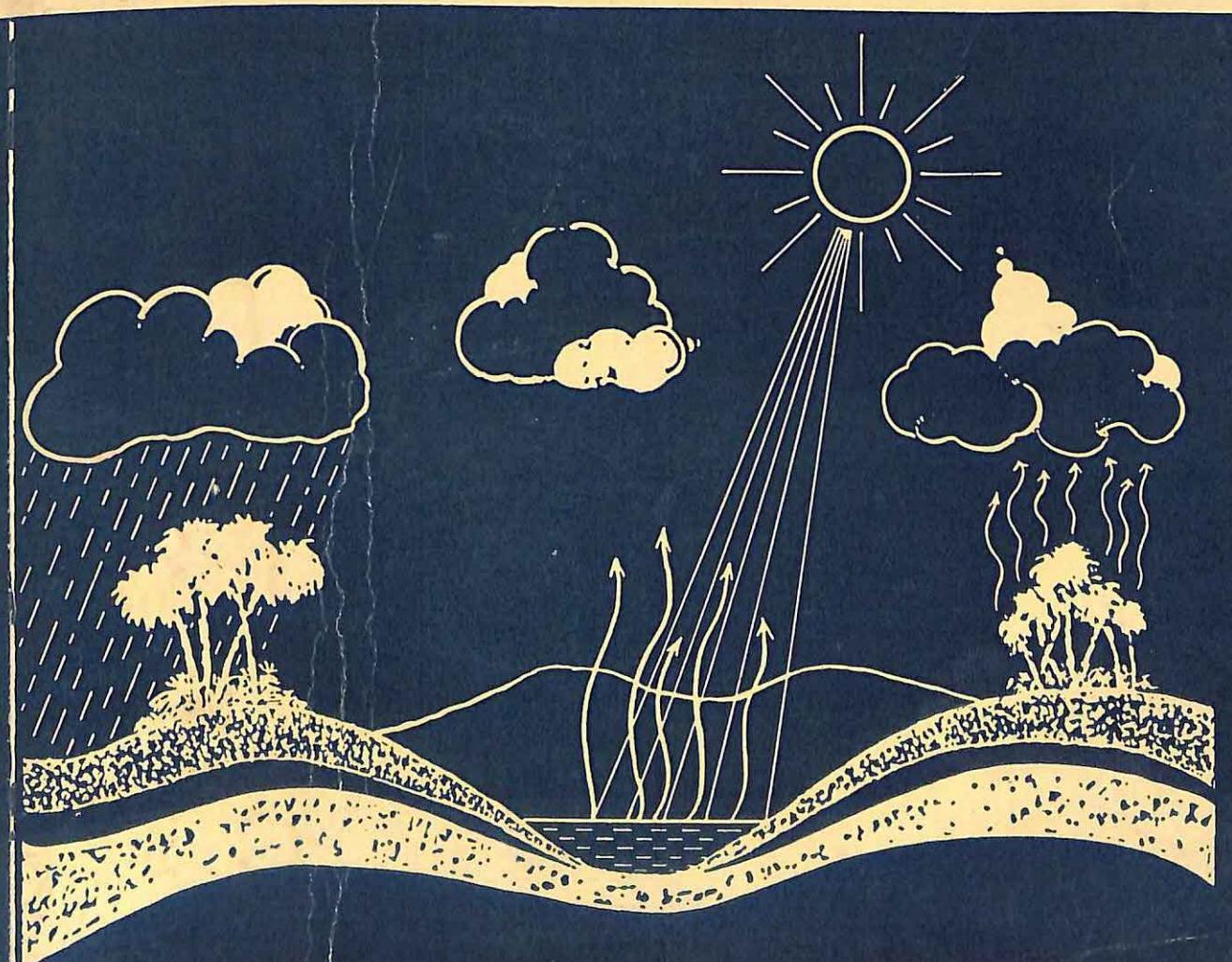


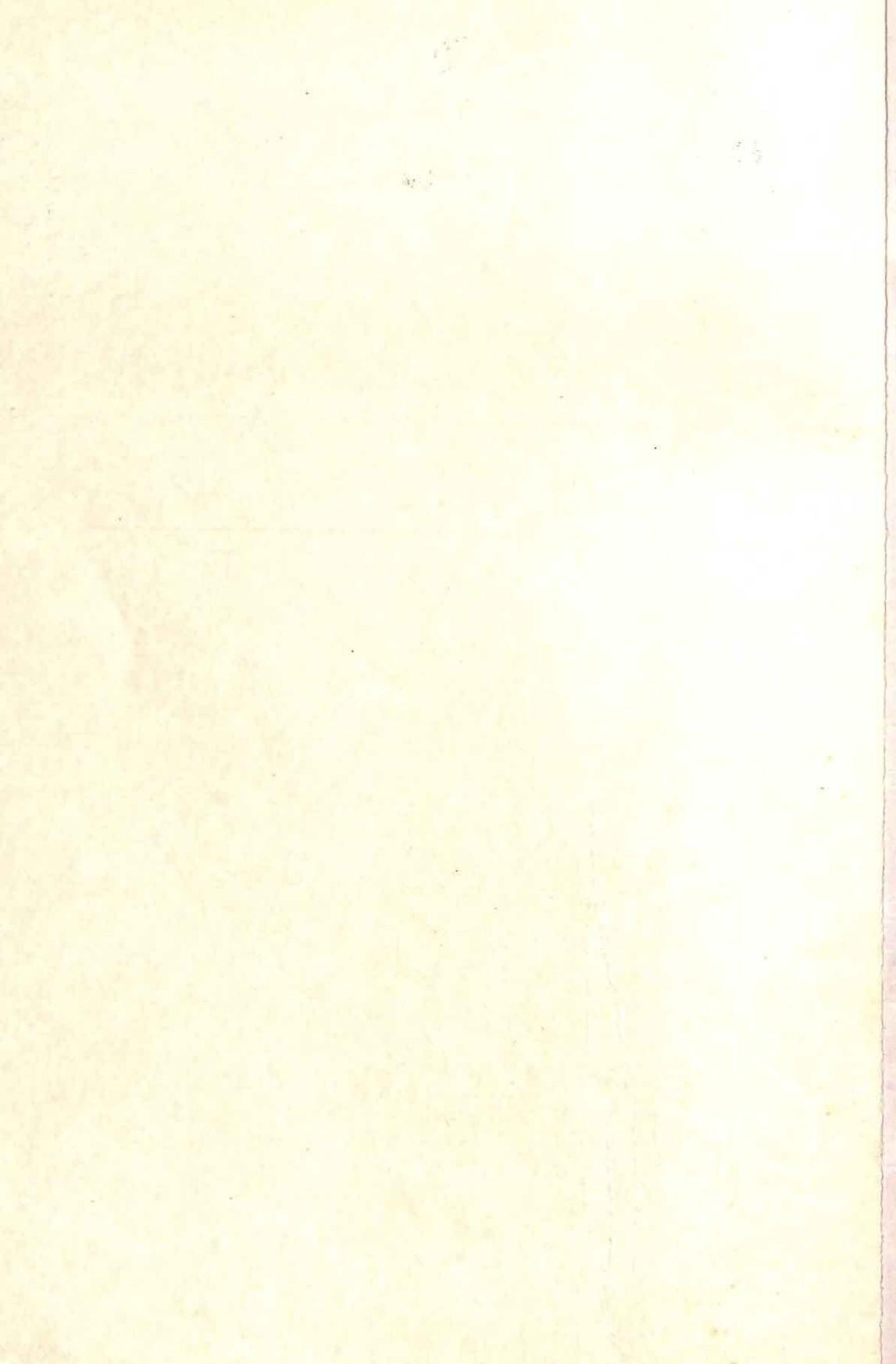
Middle School Science

CHEMISTRY

Classes 7 & 8

D.N. Verma • B.G. Pitre





Middle School Science **CHEMISTRY**

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Distributed by
Orient Longman

Middle School Science—Chemistry Classes 7 & 8

©1985 Sangam Books (India) Pvt. Ltd.

First published 1985

Reprinted 1985, 1986, 1987, 1988, 1989, 1990

ISBN 0 86131 525 1

Published by Sangam Books (India) Pvt. Ltd.

3-5-820 Hyderguda, Hyderabad 500 029 (A.P.)

Distributed by

Orient Longman Ltd.

Bombay Calcutta Madras New Delhi

Bangalore Hyderabad Patna Guwahati Lucknow

ACC NO - 15348

Cover design by J M S Rawat



Design & Layout : Rajesh Gupta

Printed at :

Gopsons Papers Pvt. Ltd., A-28, Sector IX, NOIDA

Preface

Chemistry is generally considered to be a dull subject in which only facts are to be memorised. This leads to the students losing interest in it at a very young age. However, with proper presentation of the same facts, learning of Chemistry can become an enjoyable experience for the students.

This book is an attempt in this direction. The students approach the study of the subject through experiments designed to awaken the spirit of investigation. The aim is to encourage lively enquiry, understanding and the ability to interpret evidence. Care has been taken to present the subject in simple language and to relate the study of Chemistry to the world around us.

The sequence of topics has been carefully chosen and concepts developed gradually, taking into consideration the level of understanding of the students. No attempt has been made to introduce complicated equations until the students have assimilated the physical concepts and have understood the experimental basis of the concepts.

The first few chapters in this book are devoted to develop manipulative skills in the exploration of substances. The students are then exposed to a study of the nature of common substances in their surroundings. Finally they are led to explanations in terms of atoms, ions and molecules. The preparation and identification of some chemicals is included in the concluding chapter.

A large number of illustrative diagrams have been provided as aid to proper assimilation of the subject. The workbook portion at the end of each chapter helps the students to recapitulate and includes exercises where the student is required to apply the knowledge gained in the text.

September 1984

D N VERMA
B G PITRE

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The Scope of Chemistry

You are about to begin the study of Chemistry, a fascinating science. The chemist is keen to know about one special aspect of nature—the behaviour of various substances. He tries to find out what substances are made of, the changes which they undergo when subjected to varying conditions of temperature and pressure, and when they are brought in contact with other materials. The chemist creates new substances from old, and with these he helps to build a new world. The foundation of chemistry consists of countless experimentally observed facts.

Perhaps the first person to put any chemical reaction to use was the primitive man when he first used fire to cook his food and to warm himself. He then found that the juices of berries could colour his garments and that the skins of animals could be softened and preserved. As centuries passed, man learned that certain rocks could liberate the metals copper and tin, and that the mixture of the two metals would produce bronze, the first alloy known to mankind. This was the beginning of the *Bronze Age*. In China a high degree of skill was developed in the making of ceramics in very early times. Schools of medicine existed in India as

long back as 500 B.C., during the time of Buddha.

It is impossible to say when Chemistry began. However, the modern approach to natural processes was solidly established during the eighteenth century. Modern chemistry is concerned with a wide range of subjects.

1.1 Some Applications of Chemistry

1. Food and nutrition: During the twentieth century, agriculture has made great progress with the investigation of new methods of farming. Chemists have produced chemicals called fertilisers which can double or treble the yield of crops. They have developed insecticides to protect the crops from harmful insects. Vitamins, proteins, fats and carbohydrates are all chemicals essential for growth, maintenance and repair of body tissues.

2. Shelter: The cement, iron and the paints we use for our houses are the products of the work of chemists.

3. Medicine: The human body is a complicated chemical factory. When you fall ill, medicines

are given to cure you. Various drugs, pills and injections have been prepared which are all chemicals. Attempts are being made to find drugs which can cure deadly diseases such as cancer and heart diseases.

4. Chemicals from nature: Metals are present in their ores, and come from the earth's crust. Most colouring matter are extracted from plants. Latex, obtained from rubber trees, when heated with sulphur produces the strong elastic material called natural rubber. Plants have been used by man to cure diseases. Research is continuously going on to discover new substances of medicinal value from plants. Penicillin, an important medicinal substance, was obtained from plants by Alexander Fleming. These are some of the chemicals obtained from substances found in nature.

5. Man-made fibres: Viscose rayon and cellulose acetate were the first man-made fibres for the textile industry. During the past thirty years several new synthetic fibres, such as nylon, terylene, orlon and acrilon have been prepared. These fibres are strong, crease-resistant and are not attacked by moth.

Among the several other benefits that chemistry has provided us with, are items of luxury such as cosmetics and perfumes, and fuels for cars, aeroplanes and rockets. At the same time we also feel the terrors that chemistry has added to war by providing dangerous explosives, poisonous gases, incendiary bombs, etc.

If all chemical processes were removed from our civilization, we would have no modern homes. Some understanding of the principles and applications of chemistry must therefore form a part of learning of every educated person.

1.2 The Chemistry Laboratory

The basis of all scientific knowledge is experimentation and observation. Performing experiments is not only interesting but an important part of your training.

In order that the laboratory work runs smoothly, you have to take certain *precautions*; some of these are listed below.

1. Carry out the experiment as instructed. Never perform any experiment without the permission of your teacher.
2. Inspect the apparatus for cracks or faults before use; obey the instructions for its use carefully.
3. To avoid mishaps, check the labels of chemicals on the bottles before using them.
4. Never taste chemicals; do not eat in the laboratory.
5. If you have to smell a gas, do it cautiously.
6. If any chemical falls on your hand, wash with plenty of running water.
7. Record your observations as soon as you make them.
8. After you have used an apparatus or a chemical, place it back in its correct position.
9. After completing the experiment, leave the bench tops and the apparatus clean and tidy.
10. In case of an accident, inform your teacher immediately.
11. Do not waste gas or water; use only the minimum quantity of chemicals necessary for the experiment.

EXERCISES

1. (a) Name the first metal known to mankind.....
(b) What was the first alloy made by mankind?

2. (a) How have chemists helped to solve the food problem?

.....
.....
.....
.....

(b) Name some chemicals which are being used to solve the food problem.

.....

(c) Name some chemicals which are essential for the growth of the body.....

.....
.....

(d) Why are man-made fibres increasingly used?

.....
.....

3. What precautions should you take when

(a) some chemical falls on to your hand?

.....

(b) you have to smell a gas?

.....

(c) an accident takes place in the laboratory?

.....

(d) you have finished your experiment in the laboratory?

.....

4. Fill in the blanks

(a) Schools of medicine existed in India in

(b) Juices of berries were used by the early man to

(c) Chemicals used to increase the yield of crops are called

(d) Metals are obtained from found in the earth's crust.

(e) Rubber trees produce which is hardened by adding to it.

(f) Penicillin was discovered by

2

The States of Matter

Before we study about matter and the various states it can exist in, let us study the sources of energy which we use in the laboratory.

2.1 The Bunsen Burner

A bunsen burner is the most common source of heat in the science laboratory. Gas enters the barrel, which is a metallic tube, through a jet, while the air enters through an air hole (Diagram 2.1). The supply of gas is controlled by the gas tap and the supply of air by a metallic ring which changes the size of the air hole. The gas-air mixture when lighted produces a flame at the top of the barrel.

There are two main types of flames that can be obtained from a bunsen burner.

(a) *Luminous flame*: Close the air hole completely, turn the gas full on and light the burner. The flame is yellow and smoky, and is called luminous. It gives more light but less heat and the gas does not burn completely. The flame deposits soot on the vessel being heated. The luminous flame is not suitable for heating purposes.

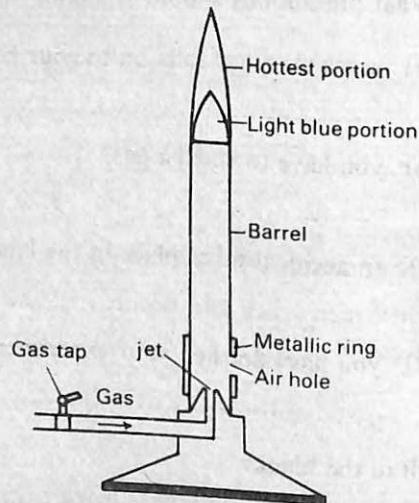


DIAGRAM 2.1 The bunsen burner.

(b) *Non-luminous flame*: When the air hole is partly or wholly open, the flame is coloured light blue and is called non-luminous. It gives less light but more heat because the gas burns completely. The flame is neither smoky nor does it deposit any soot on the article being heated. The upper zone of the non-luminous flame is the hottest part of the flame.

There is a limit to the amount of air that can be supplied through the air hole. If too much air is introduced through the air hole the flame 'strikes back', i.e. it travels down the barrel and burns the incoming gas at the jet itself. This makes the barrel hot. This phenomenon is due to the fact that the rate of burning of the gas-air mixture exceeds the speed at which the gas is going up. In such an event the supply of gas should be stopped and the barrel should be allowed to cool down. The gas can be lighted again after properly adjusting the air hole.

2.2 Electricity

Electric current is another source of energy which can be used in the laboratory for heating and to bring about chemical changes. This can be derived from accumulators which are used in cars, or from torch cells. Huge dynamos in power houses generate electricity, which is supplied to our homes and laboratories. This is used for heating, cooling and for providing light.

2.3 States of Matter

We have learnt that chemistry is the study of matter—what it is made of and the changes it undergoes. But what is matter? In scientific language *matter is defined as anything that occupies space and has mass. It can be perceived by the senses.*

Experiment 1: Place a few pieces of ice in a beaker and cover it with a watch glass. Place the beaker on a tripod stand for heating (Diagram 2.2). The ice cubes melt due to heating and change to water.

If the heating is continued for some time, the water changes into steam (evaporation), which rises up the beaker and gets cooled by coming in contact with the watch glass. The steam changes back to water by cooling (condensation) and forms droplets of water on the underside of the watch glass.

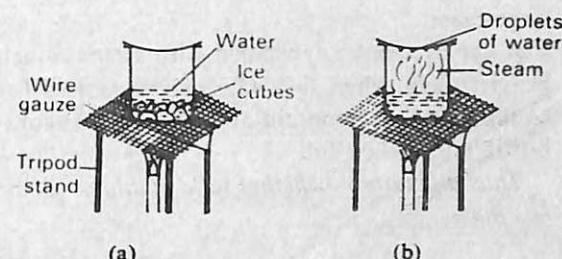
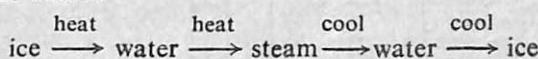


DIAGRAM 2.2 (a) The melting of ice to form water.
(b) The evaporation of water to form steam and the condensation of steam to again form water.

Now take water in a test tube and put it in a mixture of ice and common salt, called *freezing mixture*. You will notice that water changes into ice (freezing) after some time. You may also have seen water changing into ice when cooled in a refrigerator.

Ice, water and steam are composed of the same matter but are in different states, called the *solid state*, the *liquid state* and the *gaseous state* respectively. Thus there are three states of matter. Most substances can exist in all the three states under different conditions. The change of state in the above case can be shown as below:



Vapour and Gas

There is a slight difference between the terms 'vapour' and 'gas' as used by scientists. When a gas exists in the presence of the liquid of the same substance, under the same conditions of temperature and pressure, the gaseous state is called vapour. For example, when water is boiling and giving off steam, this steam is a vapour.

2.4 Properties of the Three States

1. Mass

Matter has the property of mass. A log of wood contains a lot of matter; it has a greater mass as compared to a small piece of stone which has less matter in it. It is difficult to lift a

bucket full of water because the water has a large mass.

A car or scooter tyre filled with air is much heavier than when flat. Gas cylinders used for cooking at home contain about 15 kg of combustible gas when full.

Thus all matter—whether solid, liquid or gas—has mass.

2. Volume and Shape

Experiment 2: A piece of stone, a book and a table are all examples of solids. Take a graduated cylinder. Fill it partly with water and read the volume. Tie a piece of stone with a string and lower it into the water so that the stone is fully below the water level (Diagram 2.3). Read the total volume now. You will find that this is greater than the volume of the water alone. The difference between the two readings gives the volume of the piece of stone.

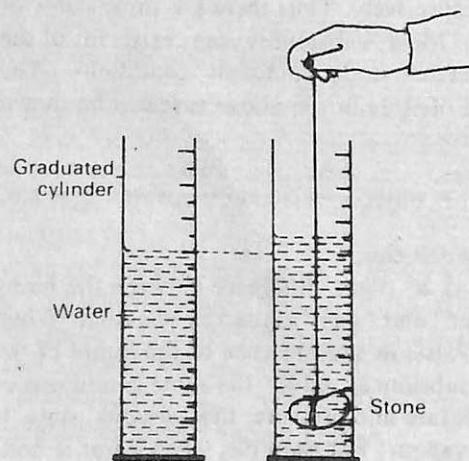


DIAGRAM 2.3 Solids have definite volume.

If the piece of stone is broken into two pieces and the volume of each piece is determined as before, it will be seen that the total volume of the two pieces is the same as the volume of the original piece of stone. The piece of stone has its own shape, which can be changed only by

the application of a force, such as by hammering. Solids thus have their own definite volume and shape.

Experiment 3: While the volume of a liquid remains the same, its shape changes according to the container it is put in. Pour some water into a graduated cylinder and read its volume. Then transfer it into a reagent bottle or flask (Diagram 2.4). You can see that the water takes the shape of the container. Pour the water back into the graduated cylinder. You will find that it has the same volume.

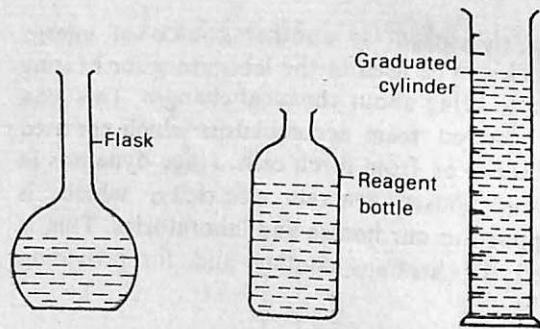


DIAGRAM 2.4 The shape of a liquid changes according to the container.

Liquids have, therefore, a definite volume but no definite shape.

Experiment 4: Take a drop of liquid bromine in a gas jar. Cover the jar with a greased glass plate and leave it for some time (Diagram 2.5a). Bromine is a volatile liquid, i.e. it has a low boiling point. It changes to a red coloured vapour at room temperature and fills the jar completely.

Now remove the glass plate and cover the jar at once with another jar, mouth to mouth, as shown in Diagram 2.5b. After some time you will notice that the red vapour of bromine fills both the jars completely.

Thus gases occupy space but they have neither a definite shape nor a definite volume.

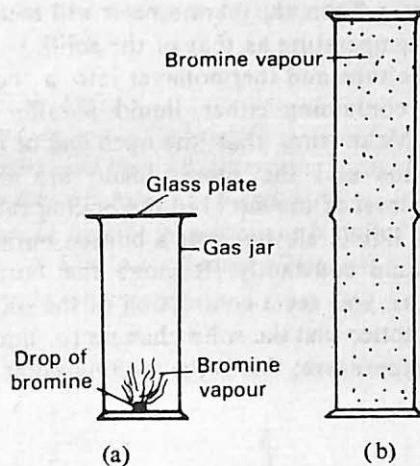


DIAGRAM 2.5 Gases have neither a definite shape nor a definite volume.

You have noticed that both gases and liquids flow. They are therefore called *fluids*.

3. Particle Nature

Take the lid off a bottle of perfume; soon its smell spreads around. Clearly, particles of the perfume spread out in the space around. A drop of liquid bromine when put into a closed jar changes to smaller particles of red coloured bromine gas which spreads in the whole jar.

Experiment 5: Put a crystal of iodine (a dark coloured solid) in a dry test tube. Fix a strip of filter paper moistened with starch solution on the mouth of the test tube with the help of a cork (Diagram 2.6). Leave the test tube for

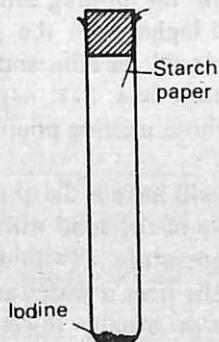


DIAGRAM 2.6

some time. You will observe the appearance of blue colour on the strip of paper.

This is due to the fact that the crystal breaks up into smaller particles of iodine, which reach the starch paper after some time, and turn the paper blue.

These experiments suggest that all the three states of matter, solid, liquid and gas are composed of small particles.

2.5 Molecules and Their Arrangement

The smallest particle of a substance that can exist independently is called a molecule. These particles are extremely small and cannot be seen by the naked eye. Molecules of any one substance are all alike and differ from the molecules of another substance.

In solids the molecules are arranged in a fixed pattern, each molecule vibrating slightly around a fixed position. The forces of attraction between molecules are very strong and free movement of molecules is not possible. Hence solids have a fixed shape.

In liquids the molecules are not held in fixed positions, but are in a state of disorder, moving around the container. These molecules are held together by weak forces of attraction so that very few break away and leave the container at normal temperature. Hence a liquid has a fixed volume but no fixed shape.

In gases the molecules move freely with great speed in all directions in straight lines until they collide with the walls of the container. There is no certainty at all of their position because the force of attraction between the molecules is very small. Hence a gas has neither a fixed shape nor a fixed volume.

Gases can mix freely with one another irrespective of their molecular masses. This is called *diffusion*.

The difference between the three states of matter as far as the arrangement of molecules is concerned is shown in Diagram 2.7.

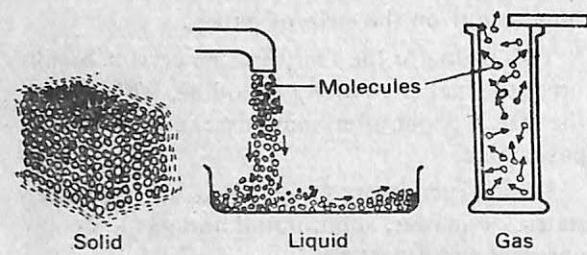


DIAGRAM 2.7 The arrangement of molecules in the three states of matter.

2.6 Pure and Impure Substances

Each pure substance has certain characteristic properties. When the substance is not pure, this property changes. A chemist knows whether a substance is pure or not by investigating these properties.

Two important properties are used to determine the purity of substances.

(a) *Melting point*: This is the temperature at which a substance changes from a solid to a liquid state.

(b) *Boiling point*: This is the temperature at which a substance changes from a liquid to a gas state.

At times it is easy to find the temperature at which a liquid changes to a solid; the process is called *solidification*. This temperature is called the *freezing point* of the substance. It is the same as the *melting point*.

Melting Point

The following method can be used to determine the melting points of naphthalene, wax or urea.

Experiment 6: Grind the solid to powder with a pestle in a mortar. Take a narrow bore tube closed at one end and fill it with about 0.5 cm layer of the powdered solid. The powdered solid is tapped down to the closed end. Attach the tube to a thermometer with a rubber band so that the solid is opposite to the bulb of the

thermometer. Then, the thermometer will record the same temperature as that of the solid.

Place the tube and thermometer into a boiling tube containing either liquid paraffin or glycerol. Make sure that the open end of the narrow tube and the rubber band are well above the level of the liquid in the boiling tube. Heat the liquid slowly over a bunsen burner. Stir the liquid constantly. Remove the burner as soon as you see a contraction of the solid. You will notice that the solid changes to liquid at this temperature; this is the melting point.

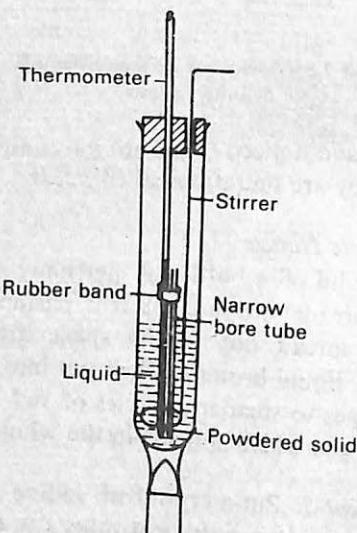


DIAGRAM 2.8 Determination of the melting point of a solid.

The liquid in the boiling tube should have its boiling point higher than the melting point of the solid. Liquid paraffin and glycerine can be used for substances like naphthalene and benzoic acid whose melting points are not very high.

A pure solid will have a sharp melting point, that is, the whole of the solid will melt at nearly the same temperature (within 0.5°C). An impure solid melts over a wide range of temperature. Impurities usually lower the melting point. The temperature at which a pure substance melts is characteristic of the solid and

hence can be used for identifying it. Pure ice melts at 0°C.

Boiling Point

The process of changing a liquid to vapour by heating and then changing the vapour back to liquid by cooling is called distillation. The liquid thus collected in a receiver is called the *distillate*. A Liebig condenser (Diagram 2.9) is used to cool the vapour.

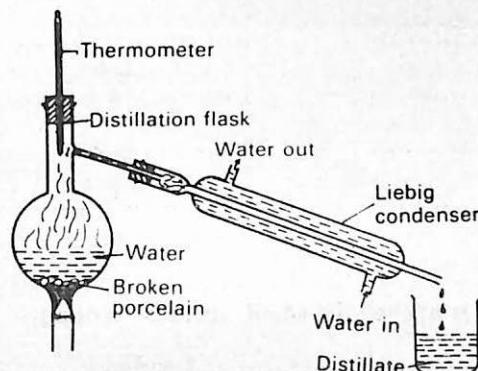


DIAGRAM 2.9 Determination of boiling point of a liquid.

The following experiment describes the method for determining the boiling point of a liquid. It can also be used to obtain the pure liquid.

Experiment 7: Take about 200 cm³ of water in a distillation flask and connect a Liebig condenser to the flask. Let a steady stream of cold water from a tap enter the lower inlet of the condenser as shown in the diagram. Fix a thermometer in the flask such that the bulb of the thermometer is near the outlet of the flask. Heat the flask and note the temperature every half minute after the water has been heated for some time. Bumping takes place when the water starts boiling. To avoid this, a few pieces of broken porcelain or porous pot should be added to the water.

After some time you will notice that in spite of continued heating the temperature remains constant, when water is continuously changing into steam. This is the boiling point of water. The pure water collected by this process of distillation is called *distilled water*. The boiling point of pure water at normal pressure (at sea level) is 100°C.

Add a weighed amount of common salt, say 5 g, to 200 cm³ of pure water and determine the boiling point again, using a sensitive thermometer. You will notice that the boiling point of the solution is more than the boiling point of pure water. Impurities raise the boiling point of a liquid.

Every liquid has its own fixed boiling point at normal pressure, a property which can be used for identifying that liquid.

EXERCISES

1. (a) What are the differences between a luminous and a non-luminous flame?

.....
.....
.....

- (b) What causes striking back in a bunsen flame?

.....
.....

(c) How can the striking back of the flame be checked?

.....
.....
.....

2. Give three differences between a solid, a liquid and a gas.

Solid

Liquid

Gas

(a)
.....
(b)
.....
(c)
.....

3. Explain the following statements with reasons.

(a) When the lid of a bottle containing a perfume is opened, its smell spreads around.....

.....
.....

(b) Solids have a fixed shape whereas liquids take the shape of the container.....

.....
.....

(c) Some solids when heated exhibit a sharp melting point, others do not.....

.....
.....

4. Fill in the blanks in the following sentences with appropriate words.

(a) Liquids and gases are called

(b) The amount of water displaced by a stone immersed in water is equal to its.....

(c) The process of a liquid changing to solid is called

(d) Freezing mixture contains and mixed together.

(e) The pressure exerted by a gas is due to of molecules of gas with the of the container.

(f) The boiling point of pure water at sea level is and melting point of ice is

Define the following terms.

(a) Molecule

(b) Volatile liquid

(c) Diffusion

(d) Matter

(e) Melting point

(f) Boiling point

(g) Distillation

6. State whether the statements given below are true or false. In the case of an incorrect statement, write the correct statement, giving reasons.

(a) A gas jar of hydrogen is put over a jar containing carbon dioxide gas (a heavy gas) mouth to mouth. Carbon dioxide stays in its own jar.....

.....
.....
.....

(b) A non-luminous flame is hotter than a luminous flame.

.....
.....

(c) A non-luminous flame leaves black deposit on the article which is heated.....

.....
.....

(d) There is no difference between the vapour and the gas state of a substance.....

.....
.....

(e) In the determination of melting point, the boiling point of the liquid in which the narrow bore tube containing the solid is immersed, should be less than the melting point of the solid.....

.....
.....
.....

Separation of Substances from a Mixture

If a bagful of coins containing 5, 10 and 25 paisa coins is spread out in front of you, it is quite easy for you to separate them into three different piles, each pile containing coins of the same value. It is equally easy to separate out marbles and stones when they are mixed together. Quite often chemists are faced with the problem of separating substances when the difference between different kinds of substances is not so distinct. The process of separation is then not so easy. Let us study some of the simple methods which chemists use in the separation of substances present in a mixture.

3.1 Mixture

A mixture can be *heterogeneous* or *homogeneous*.

A *heterogeneous mixture* is one in which the particles of the different substances can be separately seen. Take a mixture of common salt, powdered sulphur and sand. On observing the mixture through a magnifying lens you can see three different kinds of particles. The mixture is heterogeneous.

A *homogeneous mixture* is uniform through-

out, and the particles of the different substances cannot be separately seen. A homogeneous mixture is called a *solution*. Let us study its properties.

Solution

Experiment 1: Add some sugar to water in a test tube and shake well. Crystals of sugar are no longer visible. The liquid formed tastes sweet. Clearly the sugar is present in every portion of the liquid. The molecules of sugar separate from the solid, diffuse between the molecules of water and move between them.

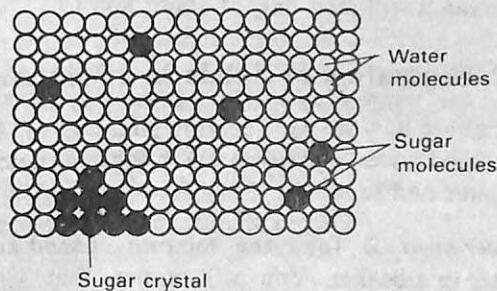


DIAGRAM 3.1 How a sugar crystal dissolves in water.

The sugar is said to be *soluble* in water; or the sugar *dissolves* in water.

The substance that does the dissolving is called the solvent. The substance that gets dissolved is called the solute. The homogeneous mixture of the solute and the solvent is called the solution.

Thus sugar is the solute and water is the solvent and both of them form a solution of sugar and water.

When a solution is heated, the solvent is removed as vapour, leaving the solute behind. This is called *evaporation*.

Liquids other than water may act as solvents. Iodine, sulphur and phosphorus are not soluble in water, but they all dissolve in carbon disulphide. Carbon tetrachloride also dissolves sulphur. Chloroform can dissolve iodine. Grease dissolves in petrol. Hence this solvent is used for dry cleaning. Paints and oils dissolve in turpentine oil.

Sulphur, phosphorus and iodine do not dissolve in water and are said to be *insoluble* in water.

Melting and Dissolving

The terms melting and dissolving are often confused. A solid melts by itself when it is heated till it changes to a liquid. A solid will dissolve with the help of another substance called the solvent.

A solid must be heated in order to melt it. Dissolving may take place without heating. Common salt dissolves readily in cold water whereas it will melt only at about 800°C .

3.2 Separation of Solids in a Mixture

Let us see how we can separate mixtures of (i) sand and water, and (ii) common salt, powdered sulphur and sand.

Experiment 2: Take the mixture of sand and water in a beaker. You will notice that sand settles down at the bottom (Diagram 3.2a) forming a *sediment*. Pick up the beaker and

very carefully pour out the water into another beaker with the help of a glass rod (Diagram 3.2b). Take care not to disturb the sand at the bottom. This process is known as *decantation*. It is used to separate a mixture of a liquid and a solid which is insoluble in it. The solid should be heavier than the liquid.

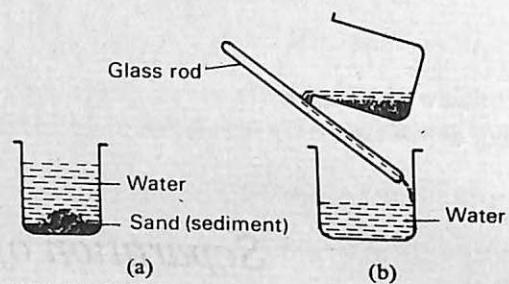


DIAGRAM 3.2(a) Sand forms a sediment in water.
(b) Decantation.

Experiment 3: Put about 5 g of a mixture of common salt, powdered sulphur and sand into a beaker. Half fill the beaker with water, heat and stir well. Now fold a *filter paper* twice to form a cone (Diagram 3.3a). Fit the filter paper into a funnel and wet it with a few drops of water so that the filter paper sticks to the funnel. Now, using a glass rod, pour the liquid through the filter paper. Take care that the liquid does not fill more than three-fourth of the height of the filter paper. This is to avoid

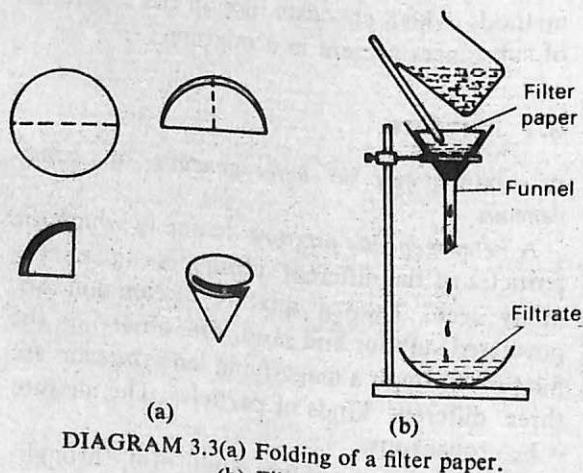


DIAGRAM 3.3(a) Folding of a filter paper.
(b) Filtration.

the flow of solid particles along with the falling liquid.

The liquid that passes through the filter paper is called the *filtrate* (Diagram 3.3b). The filtrate is a solution of common salt in water.

Sand and sulphur remain on the filter paper along with some salt which has not dissolved. Wash two or three times with hot water to free the mixture completely from common salt and collect the filtrate in the same container.

Heat all the salt solution in an evaporating dish on a tripod and gauze (Diagram 3.4). When the solution is almost dry, the hot solution and solid particles jump out of the dish; hence use a small flame towards the end. Continue heating until the water evaporates completely and dry salt is obtained. Most of the common salt we consume in our food is obtained by drying of sea water in shallow plots near the sea shore.

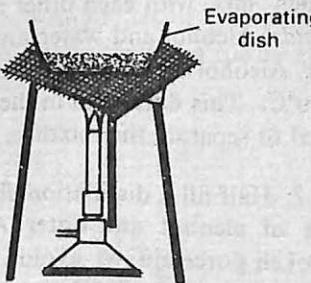


DIAGRAM 3.4 Evaporation.

Dry the mixture of sand and sulphur left on the filter paper. Take the mixture in a beaker and add some carbon disulphide and stir. The sulphur dissolves in carbon disulphide. Filter in an evaporating dish. Sand is left on the filter paper and can be collected separately. The filtrate, which is a solution of sulphur in carbon disulphide, is allowed to evaporate at room temperature. Carbon disulphide is a volatile liquid and will evaporate soon, leaving yellow particles of sulphur.

If a microscope is used to observe the sulphur, its small particles will be seen to have a definite shape with sharp edges. Such regular

formation of a solid is called a *crystal*. A crystal of sulphur has eight smooth faces and hence it is called octahedral sulphur (Diagram 3.5).

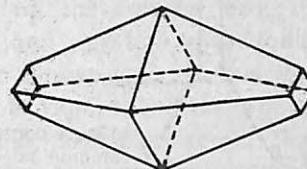


DIAGRAM 3.5 Crystal of sulphur.

3.3 Separation by Sublimation

Experiment 4: Put a small amount of ammonium chloride in a dry test tube and heat. You will observe that the solid changes directly into gas when heated. When the fumes go up in the test tube they change back to solid on cooling, without changing to the intermediate liquid state.

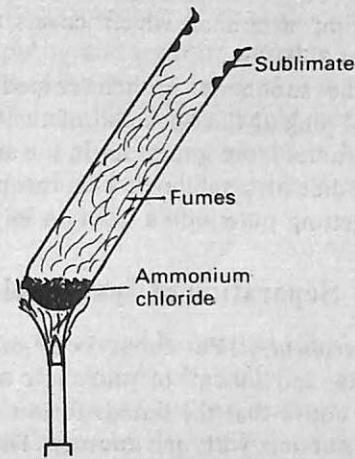


DIAGRAM 3.6 Sublimation.

Such substances are said to undergo *sublimation* and the solid deposited on the cooler part of the test tube is called a *sublimate*.

Experiment 5: When such a substance is mixed with other substances which do not sublime, separation by heating is possible.

If common salt is mixed with ammonium chloride and the mixture is heated in an evaporating dish, ammonium chloride will sublime and can be collected as a sublimate on the inner

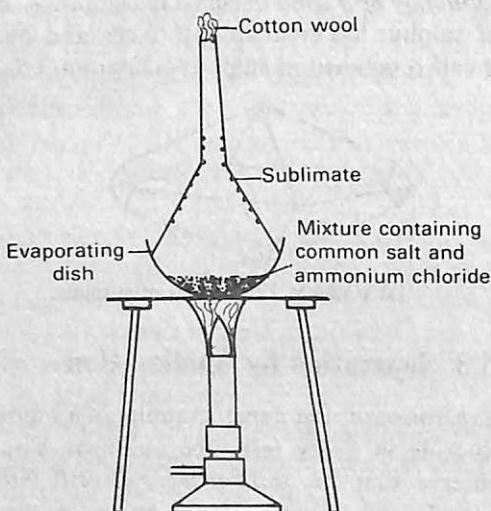


DIAGRAM 3.7 Separation of common salt and ammonium chloride by sublimation.

wall of a funnel which covers the dish (Diagram 3.7).

The sublimate is then scraped off. A cotton wool plug at the open end of the funnel checks the fumes from going out in the air.

Iodine also sublimes. This property is used in getting pure iodine from an impure sample.

3.4 Separation of Immiscible Liquids

Experiment 6: Put about 5 cm^3 of carbon disulphide and 20 cm^3 of water into a beaker. You will notice that the liquids form two layers and do not mix with one another. They are said to be *immiscible*. The more dense carbon disulphide forms the lower layer.

Transfer the two liquids into a *separating funnel*. Hold the funnel vertically, remove the stopper and run the heavier carbon disulphide into a receiver (Diagram 3.8). Stop the flow of the liquid when the line dividing the two liquids reaches the hole of the tap. Change the receiver and collect the water left into another receiver.

Kerosene oil is also immiscible with water and forms the upper layer in the separating funnel. Thus water and kerosene oil can also be separated by this method.

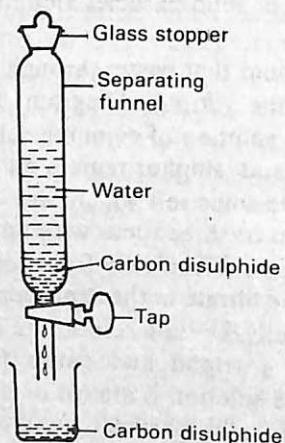


DIAGRAM 3.8 Separation of immiscible liquids using a separating funnel.

3.5 Separation of Miscible Liquids

Liquids which mix with each other are called *miscible liquids*. Alcohol and water are completely miscible. Alcohol boils at 78.5°C and water boils at 100°C . This difference in their boiling points is used to separate the mixture.

Experiment 7: Half fill a distillation flask with the mixture of alcohol and water. Add a few pieces of broken porcelain to avoid bumping. Heat the mixture to about 80°C and collect the distillate. This contains all the alcohol in the mixture and a little water which also vaporises to some extent. Take out the water left in the flask. Fill it with the distillate and distil again. The distillate will now be richer in alcohol (Diagram 3.9).

If this process were repeated several times the distillate finally obtained at about 80°C will be a sample of alcohol containing very little water. If any solid is dissolved in the mixture of liquids, it is left behind in the distillation flask.

3.6 Fractional Distillation

Collecting the distillate and re-distilling several

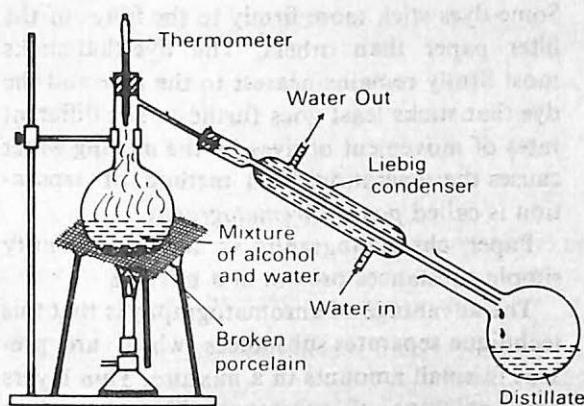


DIAGRAM 3.9 Separation of water and alcohol by distillation.

times takes a long time and the final distillate is very small in quantity. This can be avoided by using a *fractionating column* which gives the effect of several distillations (Diagram 3.10). There are many types of fractionating columns of which one is shown in the diagram.

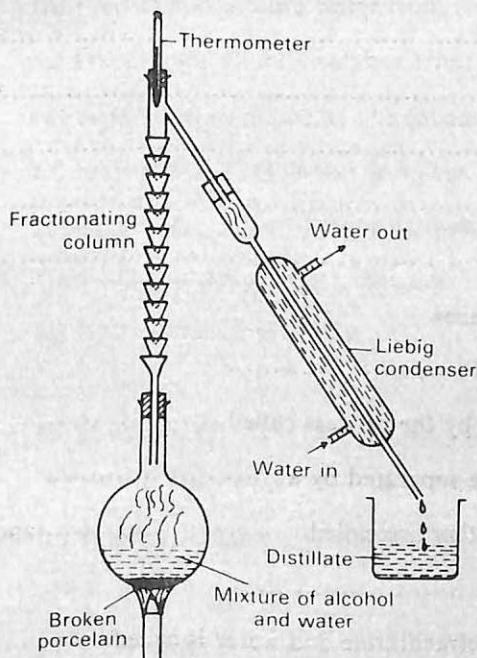


DIAGRAM 3.10 Fractional distillation.

When the mixture is heated the vapour rises up the column, cools and condenses to liquid, and falls back into the flask. The hot vapour going up the column meets the liquid falling down, and a portion of the falling liquid changes to vapour again. In this way the change of liquid to vapour, and the vapour changing back to liquid, takes place several times in the fractionating column. This gives the effect of several distillations and the vapour which finally enters the condenser contains the liquid with the lowest boiling point. Thus almost pure alcohol can be obtained from a mixture of water and alcohol.

The separation of two or more miscible liquids by distillation, making use of the difference in their boiling points is called fractional distillation. The different liquids are collected in separate receivers at their respective boiling points.

Fractional distillation is widely used in industry to purify and separate miscible liquids. More complicated columns are used in industry. Crude petroleum, a black oily liquid obtained from the earth, gives several products, known as fractions, such as petrol, kerosene oil, diesel, fuel oil and lubricating oil.

3.7 Chromatography

Chromatography means 'colour writing'. This process was developed originally to separate coloured substances present in a mixture. Nowadays chromatography is applied to separate and identify colourless as well as coloured substances.

Although there are many methods of separation of substances by this technique, a simple method of separation of different dyes in ink using paper is described below.

Experiment 8: Make a hole in the centre of a filter paper. Roll a strip of another filter paper and put it into the hole to form a wick. Put a very small spot of black ink near the hole and leave it to dry. Half fill a *petri dish* or an evaporating dish with water. Place the filter paper

on the dish (Diagram 3.11). Leave the paper for about half an hour. You will see bands of different dyes (colouring matter) present in the ink at different distances from the centre.

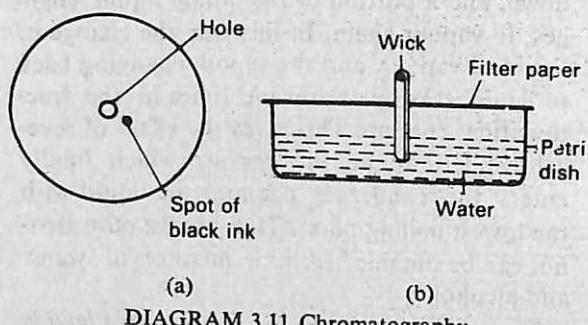


DIAGRAM 3.11 Chromatography.

The water rises up the wick, spreads across

the filter paper and carries the dyes with it. Some dyes stick more firmly to the fibres of the filter paper than others. The dye that sticks most firmly remains nearest to the hole and the dye that sticks least goes furthest. The different rates of movement of dyes by the moving water causes the separation. This method of separation is called *paper chromatography*.

Paper chromatography is used to identify simple substances present in a mixture.

The advantage of chromatography is that this technique separates substances which are present in small amounts in a mixture. Thin layers and columns of substances like aluminium oxide and silica gel are also used instead of paper to separate substances by chromatography.

EXERCISES

1. Define the following terms, giving two examples for each.

(a) Solvent.....

(b) Solute.....

(c) Sublimation.....

2. Fill in the missing words in the following sentences.

(a) Solids which have a regular shape are called.....

(b) A sediment can be removed from the liquid by the process called.....

(c) If iron is mixed with another solid, it can be separated by a.....

(d) Liquids which mix completely with one another are called..... and the apparatus used to separate them is called.....

(e) The apparatus used for separating carbon tetrachloride and water is called..... because these liquids are.....

3. (a) What do you understand by fractional distillation? Give two applications of this technique in industry.....
.....
.....
.....

- (b) How would you separate sulphur, charcoal and potassium nitrate from a mixture of these substances?.....
.....
.....
.....

4. Here is a list of chemical processes: chromatography, sublimation, distillation, crystallisation, fractional distillation, evaporation.
Which of the above is used in the following examples.

- (a) Preparation of pure iodine from impure iodine.....
(b) Preparation of pure water from sea water.....
(c) Preparation of solid sulphur from its solution in a liquid.....
(d) Separation of pigments present in a sample of ink.....
(e) Preparation of table salt from sea water.....
(f) Separation of petrol from kerosene oil.....

5. How will you remove the following?

- (a) Grass stains from a cloth.....
.....
(b) Greasy spots from your woollen garments.....
.....
(c) Sulphur sticking to a crucible.....
.....
(d) Salty water from your body.....
.....

6. (a) State how different substances form separate bands in paper chromatography.....

.....
.....
.....

(b) What is the difference between melting and dissolving?.....

.....
.....
.....

7. (a) A solution of common salt and a sample of muddy water are placed separately in two beakers. Which one is heterogeneous and which is homogeneous?.....

.....
.....
.....
.....

(b) How does a fractionating column give the effect of several distillations?.....

.....
.....
.....

(c) How would you separate ammonium chloride when it is mixed with potassium chloride?

.....
.....
.....

The Air Around Us

The Earth is surrounded by a layer of air, called the *atmosphere*, which extends to a height (altitude) of about 80 kilometres. Although it cannot be seen, its presence is felt when wind blows. Kites and birds fly as high as air takes them and the movement of clouds is directed by the air.

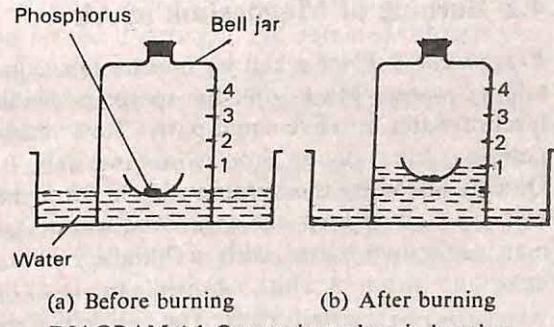
A very important use of air to us is in breathing. Without air we cannot survive even for a few minutes. There is no air on the moon and astronauts have to carry their own supply of air when they land on the moon. Mountaineers, when they climb Mt. Everest, have to carry their supply of air because at a high altitude the supply of air is not sufficient for normal life. Air dissolved in water is essential for the survival of fish. Man-made satellites must be made to orbit at altitudes of over 300 km so that they are not slowed down or burnt out by too much friction with the air.

4.1 Burning in Air

Experiment 1: Light a candle and place it on the table. Cover it with a jar. You will notice that it blows out within a few minutes. Thus burning of any substance requires the presence

of air. Let us study which part of air is used up in burning.

Experiment 2: Float an evaporating dish containing white *phosphorus* in a trough of water and place a bell jar over it. Divide the space above the level of water into five equal parts (Diagram 4.1a). Burn the phosphorus by touching it with a hot iron wire. Withdraw the wire and quickly put the stopper on the bell jar. The phosphorus burns with a yellow flame giving off dense white fumes.



(a) Before burning (b) After burning
 DIAGRAM 4.1 Oxygen is used up in burning.

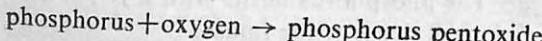
While the phosphorus is burning, you will notice that the level of water inside the bell

jar goes down slightly. The heat produced by burning of phosphorus causes expansion of air, and hence the level of water falls in the beginning.

After some time phosphorus stops burning. When the gases inside the jar cool down, the level of water inside the jar rises up to the first mark (Diagram 4.1b). Thus one-fifth of the air has been used up in burning. It is this part of the air which supports burning of substances in air. Water takes its place.

Remove the stopper of the bell jar and quickly introduce a lighted splint into it. The splint is extinguished. This shows that the remaining four-fifths of the air does not support burning. Air, therefore, contains at least two gases. The active part of the air which supports burning is called *oxygen*; and the inactive part which does not support burning contains mainly *nitrogen*.

You will also notice that the dense white fumes disappear after some time by dissolving in water. These fumes are of a new substance formed by the combination of phosphorus with oxygen of the air. The substance formed by the combination of oxygen is called an *oxide*. Hence the substance formed in this *reaction* is called phosphorus oxide, more correctly phosphorus pentoxide. The *word equation* is given below:



4.2 Burning of Magnesium in Air

Experiment 3: Place a bell jar in a trough containing water. Mark off the space above the level of water into five equal parts. Tie a magnesium ribbon to an iron wire and light it. Quickly introduce the burning ribbon into the bell jar and tighten the cork. You will notice that magnesium burns with a blinding white light and forms a white ash, while the level of water rises to the first mark. The gas left over does not support burning (Diagram 4.2).

Like phosphorus, magnesium also combines with the active part of air called oxygen, leaving

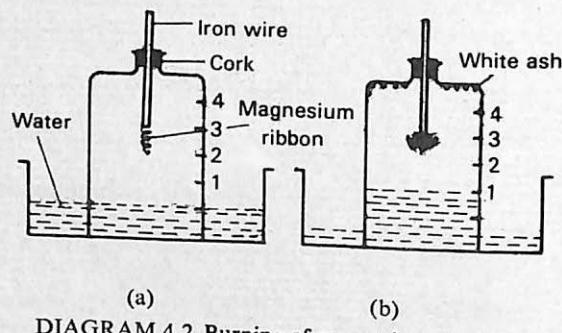
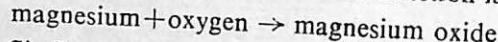


DIAGRAM 4.2 Burning of magnesium in air.

four-fifths which is inactive and is mainly nitrogen. The white ash is the new substance formed by the combination of magnesium with the oxygen of the air and is called magnesium oxide. The word equation for the reaction is:



Similar experiments with other substances show that when substances burn in air they react with the oxygen part.

4.3 Change of Mass During Burning

Experiment 4: Weigh a clean and dry crucible with the lid. Put about 10 cm long folded magnesium ribbon in the crucible. Cover it with the lid and weigh again. Remove the lid and heat the crucible strongly. As soon as the ribbon begins to burn, cover the crucible with the lid (Diagram 4.3).

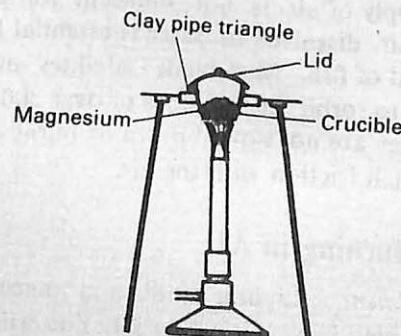


DIAGRAM 4.3 Change of mass during burning.

Raise the lid slightly so as to allow the air to enter but make sure that the white particles of magnesium oxide do not go out of the crucible. Cover the crucible with the lid and cool it in a desiccator (Diagram 9.4). This is to avoid any water vapour from the air depositing on the crucible. Weigh the crucible containing magnesium oxide with the lid. Let

$$\text{Mass of crucible with lid} = ag$$

$$\text{Mass of crucible, lid and magnesium} = bg$$

$$\text{Mass of crucible, lid and magnesium oxide}$$

$$= cg$$

$$\text{Hence, mass of magnesium} = (b-a)g$$

$$\text{Mass of magnesium oxide} = (c-a)g$$

You will notice that the magnesium oxide weighs more than the magnesium. The increase in mass is due to the mass of oxygen of the air which has combined with magnesium. Hence oxygen, and therefore air, has mass.

The products of burning in air contain a greater amount of matter than the original substance.

4.4 The Discovery of Oxygen

Burning of substances was known for a long time. The primitive man was able to cook his food by fire. He kept himself warm on cold nights and made tools and weapons using fire. But the scientific explanation of burning was first given after systematic and careful experiments carried out by Joseph Priestley (1774) and Antoine Lavoisier (1779).

Priestley performed his classic experiment in which he heated red mercury oxide by focussing sun rays on it using a magnifying lens. A modified Priestley's apparatus is shown in Diagram 4.4.

The heated red mercury oxide gave out a colourless and odourless gas which burned a candle more brightly. Priestley also observed that a mouse became more active in the gas.

Priestley told Lavoisier of his experiment. Lavoisier then carried out his classic twelve-day experiment. He heated pure mercury in a retort

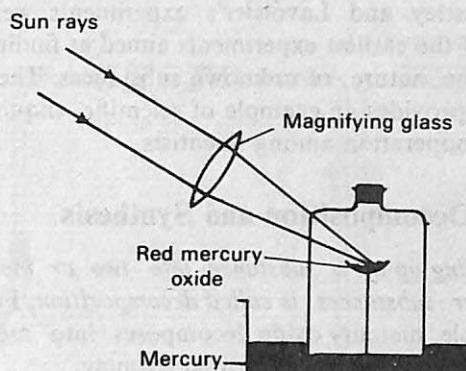


DIAGRAM 4.4 Modified Priestley experiment for preparing oxygen.

for twelve days at a temperature just below the boiling point of mercury. He noticed a red powder on the surface of the hot mercury and the level of mercury rose by one-fifth in the bell jar (Diagram 4.5).

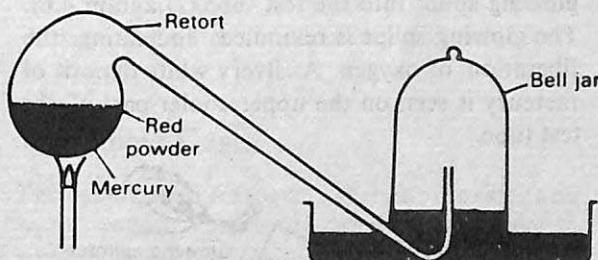
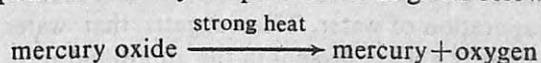


DIAGRAM 4.5 Lavoisier's twelve-day experiment.

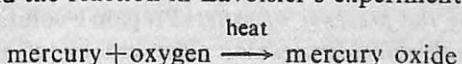
The remaining gas in the bell jar could not support life or burning.

Lavoisier collected the red powder carefully and heated it strongly. He obtained 'Priestley's gas' which caused a glowing splint to burn brightly and in which small creatures were more active.

Lavoisier named this gas oxygen. The remaining four-fifth gas which does not support burning is mainly nitrogen. The change which took place in Priestley's experiment is as given below:



And the reaction in Lavoisier's experiment was:



Priestley and Lavoisier's experiments were one of the earliest experiments aimed at finding out the nature of unknown substances. Their work provides an example of scientific enquiry and cooperation among scientists.

4.5 Decomposition and Synthesis

Breaking up of a substance into two or more simpler substances is called decomposition. For example, mercury oxide decomposes into mercury and oxygen when heated strongly.

Building up of a substance from two or more simpler substances is called synthesis.

The formation of mercury oxide when mercury is heated in oxygen is an example of synthesis.

Experiment 5: Heat some red mercury oxide in a test tube. After a few minutes, introduce a glowing splint into the test tube (Diagram 4.6). The glowing splint is rekindled, indicating the liberation of oxygen. A silvery white deposit of mercury is seen on the upper cooler part of the test tube.

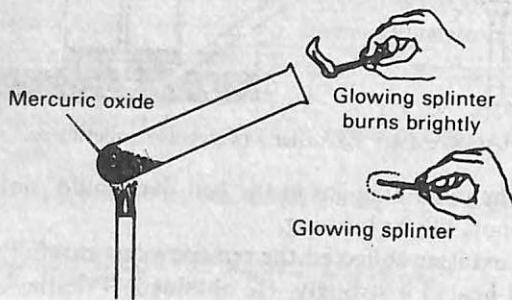


DIAGRAM 4.6 Liberation of oxygen by heating red mercury oxide.

4.6 Water Vapour in Air

The upper part of the atmosphere carries clouds which give us rain. Wet clothes dry due to evaporation of water. This suggests that water vapour is always present in the air around us.

Testing the presence of water: Prepare a solution of cobalt chloride in water. The colour of the

solution is pink. Dip a strip of filter paper in the solution and dry the strip by keeping it well over the flame of a bunsen burner. When the paper dries, the colour of the strip becomes blue. A drop of water will turn the dry strip of paper from *blue to pink*. Dry cobalt chloride paper can therefore be used for testing the presence of water.

Experiment 6: Mix a few pieces of ice with a small amount of common salt. This is called *freezing mixture* and has a temperature lower than that of pure ice. Fill a conical flask with the mixture. Allow it to stand for about 20 minutes.

You will notice that the flask gets covered with small drops of water which freeze to ice. Some of the drops drip and can be collected in a watch glass. It can be confirmed that these are drops of water by testing them with dry cobalt chloride paper. They change the colour of the paper from blue to pink.

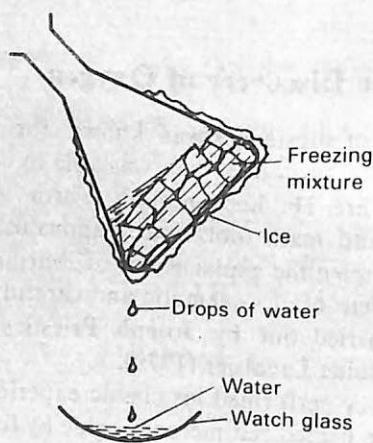


DIAGRAM 4.7 Air contains water vapour.

Clouds and fog consist of small droplets of water condensing on dust particles present in air.

4.7 Carbon Dioxide in Air

Experiment 7: Take some *lime water* in two

bottles fitted with rubber stoppers. These are fitted to an apparatus shown in Diagram 4.8. If you suck in air through the tube *T*, the air enters through Bottle 1 and when you breathe out, the air goes out through Bottle 2. Now suck in and breathe out a few times.

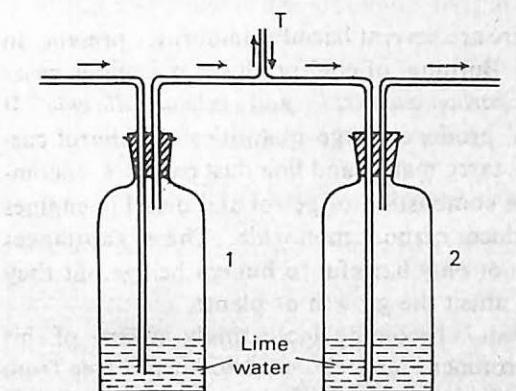


DIAGRAM 4.8 Exhaled air contains carbon dioxide.

You will observe that the lime water through which the *exhaled* air is passed forms a milky suspension. This is due to a gas called *carbon dioxide*.

Actually, some carbon dioxide is also present in the air we *inhale*. However, it is present in such a small quantity that the lime water in Bottle 1 will become milky only when air is sucked in for almost half an hour.

4.8 Products of Burning of a Candle

Set up the apparatus as shown in Diagram 4.9. Light a candle and run the suction pump so that the gases produced by the burning of the candle are drawn through the apparatus.

You will notice that the cobalt chloride paper in the tube turns from blue to pink, showing that water vapour is one of the products of burning. The lime water in the bottle turns milky, indicating that carbon dioxide is also formed.

Energy in the form of light and heat is given out when the candle burns. Coal and oil are

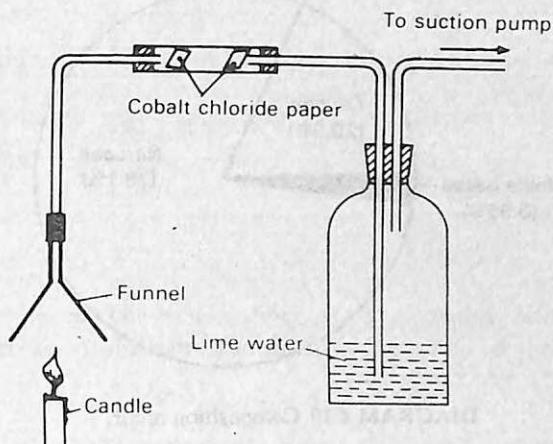


DIAGRAM 4.9 Gas produced on burning a candle contains water vapour and carbon dioxide.

burned in power stations; petrol is burned in cars and scooters. They all produce carbon dioxide and water vapour along with some form of energy.

4.9 Noble Gases

Besides oxygen, nitrogen, carbon dioxide and water vapour, there are other gases present in small quantities in the air around us. These are called *noble* or *inert gases*. They are helium, neon, argon, krypton and xenon. There is another noble gas called radon obtained by the decomposition of radium. These gases are very inactive and they hardly react with other chemicals.

Careful measurements show that the approximate composition of air, *by volume*, is as follows (Diagram 4.10).

Nitrogen	78.1%
Oxygen	20.9%
Carbon dioxide	0.03%
Noble gases	0.95%

Water vapour and impurities are present in small amounts.

The composition of air is generally the same in the atmosphere near the surface of the earth.

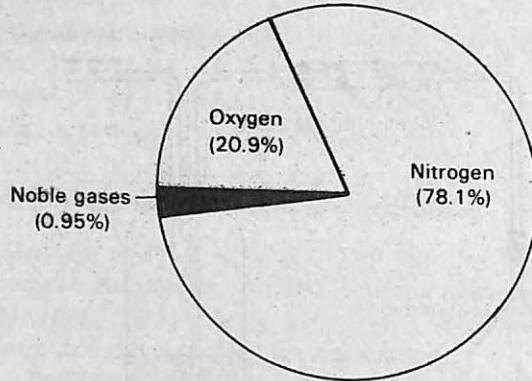


DIAGRAM 4.10 Composition of air.

The percentage of water vapour, however, varies considerably from place to place. Nitrogen, the most abundant of all the gases, is quite unreactive. One use of nitrogen is, however, quite clear. If the air contained mostly oxygen, even small fires would turn into huge explosions

and cause great damage. Nitrogen is taken in by some plants to produce *proteins*.

Carbon dioxide is required by plants for preparing substances called *carbohydrates* such as cane sugar.

4.10 Pollution of Air

There are several harmful impurities present in air. Burning of coal produces poisonous gases like *carbon monoxide* and *sulphur dioxide*. It also produces large quantities of unburnt carbon, tarry matter and fine dust particles. Incomplete combustion of petrol and diesel in engines produces carbon monoxide. These substances are not only harmful to human beings but they also affect the growth of plants.

Man is becoming increasingly aware of his environment and the need to keep it free from harmful impurities. Efforts are now being made to minimise air pollution.

EXERCISES

1. (a) Why do mountaineers carry oxygen cylinders when they go to higher altitudes?.....

.....

.....

.....

(b) What is the importance of air dissolved in water?.....

.....

.....

.....

(c) Why do satellites orbit approximately 300 km above the earth's surface?.....

.....

.....

.....

2. When phosphorus is burnt in a trough and covered with a bell jar, some of the observational changes taking place are as given below. Explain each of the observations.

(a) The level of water goes down first.....

.....

.....

.....

- (b) It rises finally to 1/5 of the space above the level of water.....
.....
- (c) White fumes are seen in the beginning.....
.....
- (d) The white fumes gradually disappear.....
.....
3. (a) A given mass of coke is burnt in a closed space. The gases formed and the residue left are all weighed. Will there be an increase or a decrease in mass as compared to the coke initially taken?.....
- (b) Give reasons for your answer stated above.....
.....
- (c) Why are chemicals sometimes put in a desiccator?.....
.....
4. Fill in the blanks in the sentences below.
- (a) When elements burn in air, they generally combine with.....of the air, and form their.....
- (b) Helium, neon, argon are called noble gases because they areelements.
- (c) Nitrogen from the air is taken by some plants to produce.....
- (d) Plants take in carbon dioxide of the air to give.....
- (e) Magnesium burns in air to form.....
5. Define the following terms.
- (a) Synthesis.....
.....
- (b) Thermal decomposition.....
.....
6. (a) How will you detect the presence of carbon dioxide in air ?.....
.....

(b) How will you find out if a given sample of water is pure?.....

.....
.....

(c) How will you find out if a given sample of a colourless liquid contains water?.....

.....
.....

(d) Name substances which cause pollution.....

.....

7. (a) Give the approximate composition of air by volume.....

.....

(b) What is the function of nitrogen in air?.....

.....

(c) Describe Lavoisier's experiment on the discovery of oxygen.....

.....
.....
.....
.....

8. (a) What is the difference between inhaled and exhaled air?.....

.....

(b) How will you distinguish between oxygen and carbon dioxide?.....

.....
.....

(c) A white crystalline solid is heated in a test tube, a colourless liquid is deposited on the upper part of the test tube. This liquid is put on a paper coated with a substance Y which turns from blue to pink.

(i) Name the colourless liquid.....

(ii) Name the substance Y.....

Elements, Compounds and Mixtures

5.1 Physical and Chemical Changes

In the earlier chapters we observed different types of changes taking place in substances. Melting of ice, condensation of water from vapour or steam, dissolving of sulphur in carbon disulphide, burning of phosphorus in air and burning of magnesium in air were some of them.

Some of these changes result in the formation of completely new substances, and are accompanied by energy changes such as evolution of heat or absorption of heat. In some cases it is possible to get back the original substance and the energy changes are not large.

Melting of ice results in the formation of water which is a liquid state of the same substance; the change needs only slight heating. The water can be changed back into ice by cooling. *Dissolving* of a solute in a solvent also produces no new substance; the solid solute can be easily obtained back by *evaporation*. If we carry out weighing in the above experiments, we would find no change of weight in the substances which have undergone these changes. **Such changes are called physical changes.**

On the other hand, *burning* of phosphorus

or magnesium in air forms new substances called phosphorus pentoxide and magnesium oxide respectively, which cannot be easily changed back into the original substances. Heat and light energy are liberated in the burning of these substances. The products show a change in mass as compared to the mass of the original substances. Such changes are called *chemical changes*.

These simple characteristics of physical and chemical changes may be summarised as shown in Table 5.1.

Millions of different materials and chemicals exist in the world. Scientists are always adding to this list. It would be almost impossible to remember how they behave, look and feel unless we find ways of grouping them into distinct categories. We know all gases flow, have no fixed shape or volume and have a low density. Thus one method of grouping is according to their state—solid, liquid and gas. Such grouping of substances, based on some common properties, is called *classification*. If we study one member of a group we can tell a lot about other members of that group. Let us consider one other method of classification of matter.

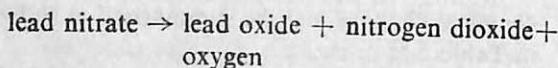
TABLE 5.1 Characteristics of physical and chemical changes.

<i>Physical changes</i>	<i>Chemical changes</i>
1. No new substance is formed (e.g. melting of ice, boiling).	1. New substances are formed (e.g. phosphorus + oxygen = phosphorus pentoxide).
2. Usually energy change is small (e.g. dissolving of sugar in water).	2. Often energy change is large (e.g. burning of magnesium in air).
3. Can be easily reversed (e.g. condensation of steam).	3. Cannot be easily reversed (e.g. decomposition of magnesium oxide).
4. There is no change of mass (e.g. melting of ice).	4. Product masses are different (e.g. burning of phosphorus in air).

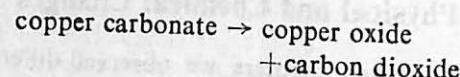
5.2 Elements

Experiment 1: You have seen that heating of mercury oxide decomposes it into simpler substances—mercury and oxygen. Heat a few crystals of lead nitrate in a test tube. The crystals break into smaller particles producing a crackling sound; this is called *decrepitation*. Oxygen is given out which rekindles a glowing splint. A reddish brown gas is also produced which is nitrogen dioxide and a yellow coloured residue is left in the test tube which is lead oxide. The reddish brown fumes when cooled change into a yellow liquid (Diagram 5.1).

Thus the reaction can be written as follows:



Experiment 2: Copper carbonate is a green powder. If a small amount of the substance is heated, it leaves a black residue of copper oxide and carbon dioxide gas is given out, which turns lime water milky (Diagram 5.2).



Both lead nitrate and copper carbonate decompose into simpler substances. Most substances can be broken down into simpler substances either by heating or by some other more complicated method.

A small number of substances, however, cannot be broken down by chemical changes into simpler substances. These substances are called elements.

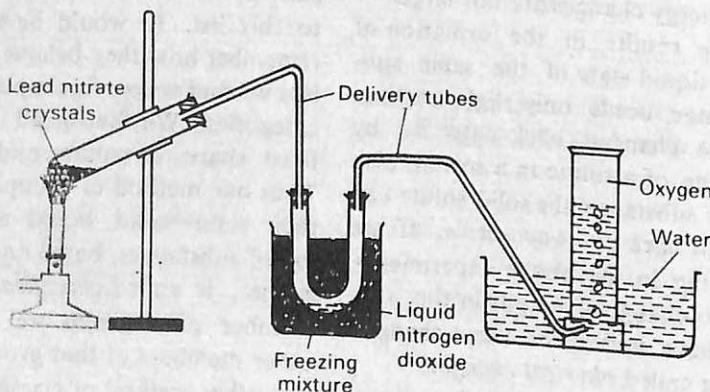


DIAGRAM 5.1 Decomposition of lead nitrate.

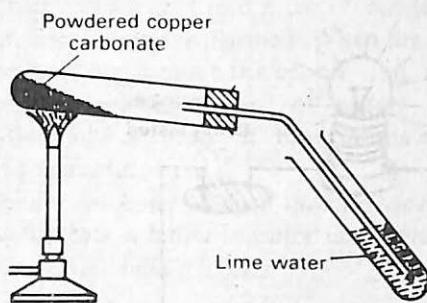


DIAGRAM 5.2 Decomposition of copper carbonate.

At present 105 elements are known to exist. Of these, 92 occur naturally in the earth and in the atmosphere, while 13 have been made by artificial means. These elements are the building block of all the millions of complicated substances. The most common elements that form a majority of substances are oxygen, silicon, aluminium, iron and hydrogen. Diagram 5.3 shows the relative abundance of nine most abundant elements in the earth's crust.

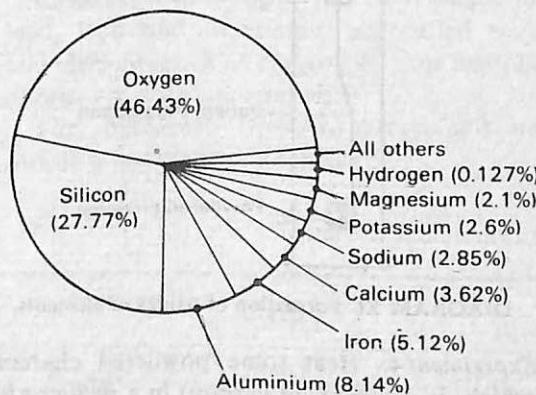


DIAGRAM 5.3 The nine most abundant elements in the earth's crust.

5.3 Properties of Matter

All substances possess certain characteristic qualities which can be used for identification of these substances. These characteristic qualities of a substance are called its *properties*.

Physical properties of a substance are those which can be determined without changing it into a new substance. Colour, density, solubility, melting point, boiling point and conduction of heat and electricity are some of the physical properties. For example, we know that water is a colourless liquid, it changes to solid state at 0°C, and its boiling point is 100°C. It is a poor conductor of heat and its density is 1 g/cm³. These are all physical properties of water.

The properties of a substance which involve chemical changes are called its chemical properties. Water reacts with phosphorus pentoxide to form a new substance. This is a chemical property of water.

Chemical properties are helpful in determining how a substance can be used to prepare other useful substances.

5.4 Classification of Elements

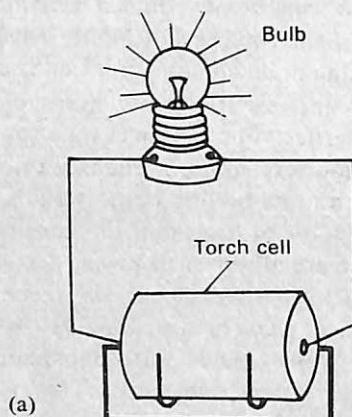
Carbon, phosphorus, sulphur, sodium, aluminium, iron, magnesium and lead are some of the common elements which should be available in your laboratory. Let us try to classify these elements according to some physical and chemical properties.

Experiment 3: Hold one end of a rod of copper, aluminium, and carbon, one by one, in your hand. Hold one end over the flame of a bunsen burner as shown in Diagram 5.4.

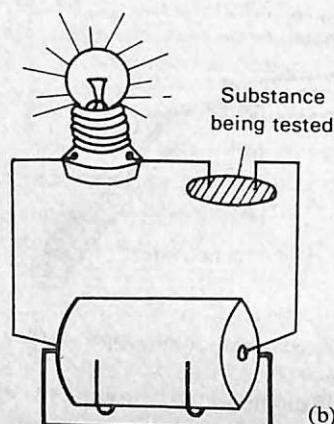


DIAGRAM 5.4 Heating a rod on a bunsen flame to determine whether it is a good or a bad conductor of heat.

You will notice that your hand feels hot very soon when you are holding the copper or



(a)



(b)

DIAGRAM 5.5 To determine whether an element is a good or a bad conductor of electricity.

aluminium rod. In case of carbon, however, it takes a very long time. Thus copper and aluminium are good conductors of heat whereas carbon is a bad conductor.

Take a torch cell and connect two wires of copper to its two ends. Join the other ends to a torch bulb; the bulb will glow (Diagram 5.5a). Now include a small piece of each element in the path of the current (Diagram 5.5b). You will find that in case of copper, aluminium and iron, the bulb glows. This shows that they are *conductors* of electricity. In case of carbon and sulphur the bulb does not glow. These are *nonconductors* of electricity. However, graphite, a form of carbon, does conduct electricity.

You will also notice that elements like phosphorus, carbon and sulphur have low *density* whereas copper, iron and lead have high density.

If you hammer a piece of carbon or sulphur it breaks into pieces. They are said to be *brittle*. Heated iron, aluminium and lead can be hammered into sheets. They are thus *malleable*. They can also be drawn into wires, and are said to be *ductile*.

We can therefore see that elements such as copper, iron, aluminium and lead are good conductors of heat and electricity, have high density and are malleable and ductile. On the other hand, elements such as carbon and sul-

phur are bad conductors of heat and electricity, have low density and are brittle. Thus, it is possible to classify the elements into two groups on the basis of these physical properties.

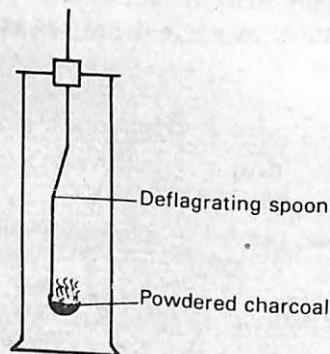


DIAGRAM 5.6 Formation of oxides of elements.

Experiment 4: Heat some powdered charcoal (which is a form of carbon) in a *deflagrating spoon* until it is red hot. Introduce it into a jar of oxygen. Charcoal burns in oxygen to form carbon dioxide. Let the charcoal remain in the jar until it stops burning. Add about 10 cm^3 of water to dissolve the carbon dioxide gas. Add to this solution a drop of *blue litmus solution*. It will turn dull red.

Take some powdered sulphur in another deflagrating spoon and heat it. When it starts

burning introduce it into a jar of oxygen. Sulphur dioxide gas is formed. When the sulphur stops burning, remove the spoon and dissolve sulphur dioxide in 10 cm³ of water. To this solution add a drop of blue litmus solution, which again turns red.

Oxides such as carbon dioxide and sulphur dioxide whose solution in water turn blue litmus red are called acidic oxides.

Experiment 5: Burn sodium in a deflagrating spoon and introduce it into a jar of oxygen. To the white residue of sodium oxide left in the spoon add some water. A colourless solution is formed. When a drop of red litmus solution is added to this, it turns blue.

When magnesium burns in oxygen it forms magnesium oxide which is dissolved in water. This solution will also turn red litmus solution blue.

Oxides such as sodium oxide and magnesium oxide whose solutions in water turn red litmus blue are called basic oxides.

Elements such as copper, sodium, magnesium, lead, iron and aluminium are called *metals*, and elements such as carbon, sulphur and phosphorus are called *non-metals*.

The difference between metals and non-metals is summarised in Table 5.2.

It must be remembered that not all metals and non-metals show all the characteristics of the class to which they belong. For example, the metal sodium has a low density and floats on water. The non-metal carbon in the form of graphite conducts electricity.

5.5 Properties of Iron and Sulphur

Let us now investigate the properties of two elements, iron which is a metal, and sulphur which is a non-metal, separately and in combination.

Experiment 6: Examine the elements iron and sulphur. You will see that iron is a grey metal and sulphur is a yellow solid.

Take some powdered sulphur in a test tube and add carbon disulphide to it. Sulphur dissolves to give a clear solution. Now add carbon disulphide to iron filings. You will find that they do not dissolve.

Bring a magnet near powdered sulphur and iron filings. You will find that iron filings are attracted by the magnet whereas sulphur is not (Diagram 5.7).

Experiment 7: Take each of the two substances (iron and sulphur) separately in two test tubes.

TABLE 5.2 Differences between metals and non-metals.

Metals	Non-metals
<ol style="list-style-type: none">Good conductors of heat and electricity.Have generally high density.Are usually shiny (lustrous).Have high melting point.Are usually malleable (can be cast into sheets) and ductile (can be drawn into wires).Usually combine with oxygen to give basic oxides.	<ol style="list-style-type: none">Bad conductors of heat and electricity. (Graphite is an exception. It conducts electricity.)Have generally low density.Are usually non-shiny.Have usually low melting point.Are usually brittle.Usually combine with oxygen to give acidic oxides.

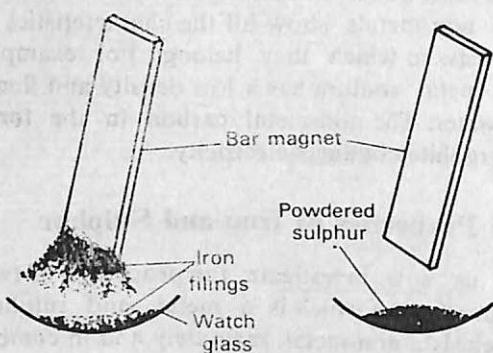


DIAGRAM 5.7 Iron filings are attracted by a magnet whereas sulphur is not.

Add *dilute hydrochloric acid* into the test tubes. You will notice that there is no reaction with sulphur. But the acid reacts with iron filings with *effervescence*, and bubbles are seen in the test tube (Diagram 5.8). This indicates the evolution of a gas. When a *lighted splint* is introduced into the gas, the splint goes off and the gas burns with a *pop* sound. This is the property of *hydrogen gas*. Thus the acid reacts with iron to give hydrogen.

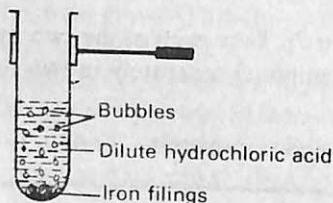


DIAGRAM 5.8 Reaction of iron filings with dilute hydrochloric acid to give hydrogen gas.

5.6 Mixture of Iron and Sulphur

Experiment 8: Mix two parts by weight of iron filings and one part by weight of powdered sulphur thoroughly on a piece of paper. Examine the *mixture* with a hand lens. You can see iron and sulphur particles separately in the mixture. When a magnet is brought close to the mixture, iron is attracted by the magnet (Diagram 5.9).

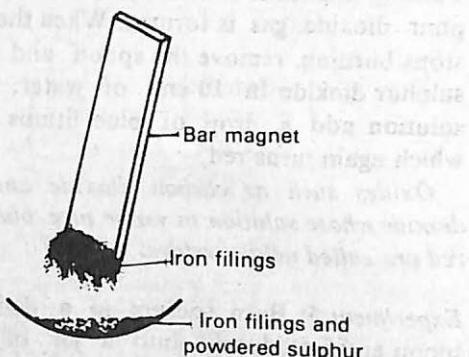


DIAGRAM 5.9 A magnet attracts iron filings from a mixture of iron filings and powdered sulphur.

Take some of the mixture in a test tube and add about 5 cm^3 of carbon disulphide. Shake well and allow the test tube to stand for some time. Sulphur dissolves in carbon disulphide leaving iron filings as the residue. Now if a few drops of the carbon disulphide solution are taken on a microscope slide, carbon disulphide evaporates, leaving crystals of sulphur which can be seen by a microscope.

Experiment 9: Take a small sample of the mixture in a test tube and add to it some dilute hydrochloric acid. You will again see effervescence. When a lighted splint is introduced into the test tube, the gas coming out burns with a 'pop' sound showing that hydrogen gas is evolved.

Thus, both iron and sulphur retain their physical and chemical properties in the mixture. No new substance is formed in the mixture. You will also notice that no change in energy takes place when iron and sulphur are mixed, i.e. heat is neither given out nor does cooling take place. Iron and sulphur can be easily separated by physical methods, like dissolving and attraction by a magnet.

5.7 Compound of Iron and Sulphur

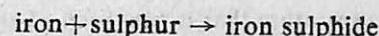
Experiment 10: Place a little of the mixture of iron filings and powdered sulphur in a test tube

and heat the test tube over a flame. When the mixture starts to glow, remove the test tube from the flame. You will notice that the glow continues even after this, indicating that a lot of heat is given out in the reaction. Such a reaction in which heat energy is given out is called an exothermic reaction. Allow the test tube to cool and collect the residue in a watch glass. On examining the residue by a hand lens you will not see the particles of iron and sulphur separately. Instead a black solid is formed. When a magnet is taken close to the residue, particles are not attracted.

When carbon disulphide is added to the residue, the solution formed does not contain sulphur particles as it does not leave particles of sulphur when evaporated.

Clearly, individual physical properties of iron and sulphur are no longer present in the residue left after heating. It appears that a new substance has been formed from which iron and sulphur cannot be separated by physical means.

Experiment 11: Take another portion of the residue left after heating in a test tube and add dilute hydrochloric acid. You will notice an effervescence showing that a gas is evolved. The gas is not odourless like hydrogen but it has an unpleasant smell like rotten eggs. The gas is *hydrogen sulphide* and will turn a strip of paper soaked in *lead acetate solution* black. This shows that when iron filings and sulphur are heated, a new substance is formed. The new substance is called *iron sulphide*. The chemical equation is written as follows:



Iron sulphide is a compound of iron and sulphur and has neither the physical nor the chemical properties of iron or sulphur.

Taking this example as a typical one, we can now summarise the difference between a mixture and a compound (Table 5.3).

TABLE 5.3 Differences between a mixture and a compound.

Mixture	Compound
<ol style="list-style-type: none"> When a mixture is prepared, very little or no energy change takes place. Formation of a mixture is a physical change. Separate particles of the constituents of a mixture can usually be seen. The constituents retain their individual properties. No new substance is formed. The constituents can be separated by physical methods. The constituents of a mixture need not be mixed in a fixed proportion. 	<ol style="list-style-type: none"> When a compound is formed an energy change usually takes place; heat is given out or absorbed. Formation of a compound is a chemical change. Separate particles of the chemicals present in a compound cannot be seen. The properties of a compound are completely different from those of its component elements. A new substance is formed. Physical methods cannot separate the component elements of a compound. The constituents of a compound are in a fixed proportion. (This fact has not been proved in the above experiments; actually in any sample of iron sulphide, iron and sulphur are present in a fixed ratio of 7 : 4 by weight.)

5.8 Is Air a Mixture or a Compound?

Once we have classified substances as mixtures and compounds on the basis of certain physical and chemical properties which are present in them, these tests can be applied to find out whether air is a mixture or a compound.

1. While the composition of air is generally constant, slight differences exist from place to place. The air at higher altitudes contains less oxygen. Water vapour present in air varies a lot; the climate is more *humid* (i.e. there is more water content in air) at places near the sea-shore.

2. Scientists have been able to cool air to such a low temperature that it has been changed to liquid state, called *liquid air*. From this the constituent elements, such as oxygen and nitrogen, can be separated by fractional distillation.

3. The properties of oxygen, carbon dioxide, water vapour, etc. are present in air. Oxygen present in air supports burning of substances. Lime water exposed for a long time to air turns milky. Water vapour present in air turns cobalt chloride paper from blue to pink.

4. If the gases present in air are mixed in the same proportion as they are found in air, there is no evolution of heat or any other form of energy. The gases mix together without any chemical change.

Hence, we can say that air is a mixture and not a compound.

5.9 Is Water a Compound or a Mixture?

You have observed earlier in an experiment that hydrogen is evolved when dilute hydrochloric acid is added to iron. When a burning splint is introduced into the test tube in which the gas is evolved, the gas burns with a 'pop' sound and you can see tiny droplets of water on the walls of the test tube. These droplets are formed by the burning of hydrogen in air, when the gas combines with oxygen to form water.

Thus water is formed by a chemical change when hydrogen burns in oxygen. These elements cannot be separated from water by physical methods, such as boiling or cooling. Both oxygen and hydrogen are gases at room temperature whereas water is a liquid. Hence the properties of the component elements are not retained when water is formed.

It has been found experimentally that in water, hydrogen and oxygen are present in a fixed proportion of 1 : 8 by mass.

These reasons show that water is a compound of hydrogen and oxygen.

EXERCISES

1. Classify the following substances into elements, mixtures and compounds.

- | | |
|-------------------|---------------------|
| (a) Air | (d) Gold |
| (b) Diamond | (e) Brass |
| (c) Sand | (f) Petroleum |

2. State whether the following statements are correct or incorrect.

- | |
|--|
| (a) A solution is a mixture. |
| (b) The properties of a compound are the sum total of the constituent elements. |

- (c) Physical changes are not easily reversible whereas chemical changes can be reversed easily.
- (d) Physical changes do not produce new substances.
3. (a) Burning of a candle made of wax involves both chemical and physical changes. Explain.....
.....
.....
- (b) What is the difference between physical and chemical properties of a substance? Give three examples of physical properties and one example of a chemical property of a substance.....
.....
.....
.....
- (c) Name two elements which are most abundant in the earth's crust.
-
4. Write all that you would observe when the following are heated.
- (a) Mercury oxide
.....
- (b) Lead nitrate
.....
.....
- (c) Copper carbonate
.....
5. Define the following terms.
- (a) Ductile and malleable
.....
- (b) Acidic oxide
.....

(c) Basic oxide.....
.....

6. (a) Give two reasons to prove that air is a mixture and two reasons to support that water is a compound.....
.....
.....
.....
.....

(b) Name a non-metal which conducts electricity.....

(c) How will you find whether a gas is hydrogen or oxygen?.....
.....
.....

7. Iron and sulphur are mixed together and divided into four portions. To each portion the following liquids are added separately. Write all that you would observe.

(a) Dilute hydrochloric acid

.....

(b) Carbon disulphide

.....

The other two portions are heated, cooled and then

(c) dilute hydrochloric acid is added.....
.....

.....

(d) carbon disulphide is added.....
.....

.....

What inference can you draw from the above experiments? Give reasons for your answer ...
.....

.....

.....

8. (a) Give two evidences, one physical and one chemical, to show that magnesium is a metal and carbon is a non-metal
-
-
-

- (b) Complete the following table.

<i>Substance</i>	<i>Name of compound formed when substance is burnt in oxygen</i>	<i>Action on litmus of solution in water of the compound formed</i>	<i>Inference about the compound formed</i>
Carbon
Sodium
Phosphorus
Hydrogen
Magnesium
Sulphur

9. Fill in the blanks in the following.

- (a) The souring of milk is a change.
- (b) Dissolving salt in water is a change.
- (c) Rusting of iron is a combination of and of air with iron.
- (d) The only metal which is liquid is

10. Give three differences between physical and chemical changes.

<i>Physical change</i>	<i>Chemical change</i>
1.
2.
3.

6

Oxygen

Priestley and Lavoisier obtained oxygen from mercury oxide. This is an expensive compound. There are several other compounds in the chemistry laboratory which contain oxygen. Some of these, such as potassium chlorate and hydrogen peroxide, can be used conveniently for obtaining the gas in the laboratory. We shall now study methods of preparing oxygen and knowing its properties.

6.1 Oxygen from Potassium Chlorate

Experiment 1: Potassium chlorate is a white crystalline solid and is soluble in water. Take a few crystals of the substance in a clean dry test tube and heat. The crystals will change to liquid. On heating to red hot temperature it decomposes. If a glowing splint is introduced into the test tube, the splint is rekindled, showing that oxygen is produced by strongly heating potassium chlorate.

Take a small amount of manganese dioxide, a black solid, and heat it strongly in a test tube. When a glowing splint is introduced into the test tube it is not rekindled. Hence no oxygen is given off.

Mix 4 parts by mass of potassium chlorate and 1 part by mass of manganese dioxide. Take the mixture in a dry test tube and heat gently. On introducing a glowing splint into the test tube you will notice that the splint is rekindled. The experiment shows that oxygen is given off from a mixture of potassium chlorate and manganese dioxide at a much lower temperature than if potassium chlorate were heated alone. Black particles of manganese dioxide are still visible in the test tube. We shall see later that manganese dioxide has not suffered any change.

6.2 Preparation of Oxygen in the Laboratory

Experiment 2: The method of heating a mixture of potassium chlorate and manganese dioxide can be used for the preparation of oxygen in the laboratory. The experiment should be conducted carefully. It can be dangerous if heating is not done gently and cautiously.

Mix thoroughly about 16 g of potassium chlorate and 4 g of manganese dioxide in a mortar using a pestle. Place the mixture in a hard

glass test tube and set up the apparatus as shown in Diagram 6.1.

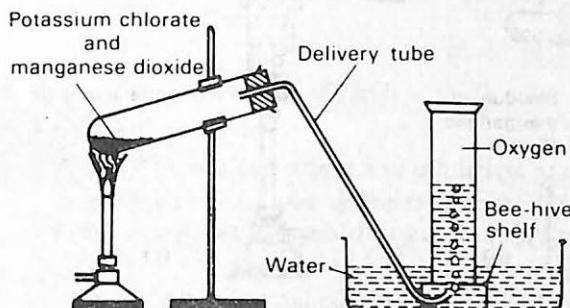


DIAGRAM 6.1 Preparation of oxygen in the laboratory.

Heat the test tube gently for about two minutes to displace the air inside the test tube and then put a gas jar filled with water over the *bee-hive shelf*. The gas jar is filled to the brim with water, covered with a well-greased lid, and lowered upside down into the trough containing water before the lid is removed. There should be no air bubbles inside the water filled jar.

Continue heating and you will see the bubbles of oxygen gas rising up and displacing the water in the gas jar. This method of collecting the gas is by *downward displacement of water*. Fill several such gas jars with oxygen.

You will notice that the gas jars containing oxygen are misty. This is due to the water vapour present along with oxygen. The hot oxygen gas vaporises a little water.

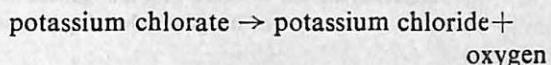
When heating is stopped, the pressure of the gas inside the test tube would become less due to cooling, and water would enter the test tube if the end of the delivery tube were allowed to remain in water. Hence the end of the delivery tube should be removed from the trough before heating is stopped.

Introduce a glowing splint into one of the jars. The splint bursts into flame. This is a good method of testing oxygen gas.

The purpose of manganese dioxide in the above experiment is to speed up the reaction. Several other substances are used to increase the

rate of reactions. In some cases a substance may also decrease the rate of a certain reaction. All such substances are called *catalysts*. A catalyst is a substance which changes the speed of a chemical reaction, but itself remains unchanged chemically and in mass. Since there is no change in the mass of a catalyst, it is effective in small quantities.

Actually, in the experiment potassium chlorate decomposes to form potassium chloride and oxygen. The chemical change can be written as follows:



6.3 Testing a Catalyst

Experiment 3: Mix 20 g of potassium chlorate with 5 g of manganese dioxide and heat gently in a hard glass test tube for about ten minutes (Diagram 6.2a). By introducing a glowing splint into the test tube you can see that oxygen gas is given off. Allow the test tube to cool and add about 15 cm³ of distilled water in it. Weigh a filter paper and fix it into a funnel. Shake well the contents of the test tube and filter through the filter paper (Diagram 6.2b).

You will notice that a black powder is left on the filter paper. Take about 5 cm³ of distilled water into the test tube to transfer any residue left. Repeat if necessary. Wash the black powder four or five times using small amounts of hot distilled water to remove the soluble potassium chloride.

Remove the filter paper and dry it in an oven. When fully dry, weigh the filter paper again. The difference between this weight and the weight of the filter paper earlier taken will give the weight of the black residue.

You will find that the weight of the residue is the same as the weight of the manganese dioxide originally taken. Its colour is also the same. To show that the black residue is manganese dioxide, add about 1 cm³ of concentrated hydrochloric acid to the residue in a test tube and

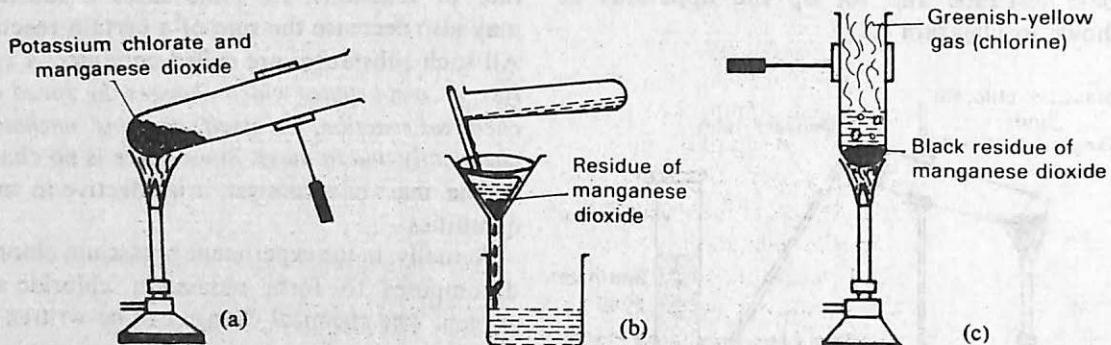
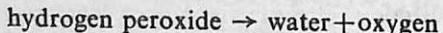


DIAGRAM 6.2 To show that manganese dioxide does not change chemically in the preparation of oxygen by heating potassium chlorate and manganese dioxide.

heat gently. A greenish-yellow gas, chlorine, is given off (Diagram 6.2c), which turns moist litmus paper colourless. This is a chemical property of manganese dioxide. This experiment shows that the black residue is manganese dioxide which has not changed chemically in the preparation of oxygen by potassium chlorate.

6.4 Oxygen from Hydrogen Peroxide

Experiment 4: Place about 2 g of manganese dioxide in a test tube and to this add about 5 cm³ of a solution of hydrogen peroxide. You will see brisk effervescence taking place. Introduce a glowing splint into the test tube. The splint is rekindled. Thus oxygen gas is evolved in the reaction. Manganese dioxide remains unchanged and acts as a catalyst. Hydrogen peroxide decomposes into water and oxygen:



This reaction is a more convenient method of preparing oxygen in the laboratory. No heating is required and the method is absolutely safe.

Experiment 5: Put about 10 g of manganese dioxide in a flat-bottomed conical flask which is fitted with a two-holed rubber stopper. The stopper carries a *dropping funnel* and a bent delivery tube. Open the tap of the dropping funnel and run in a small quantity of hydrogen

peroxide at a time. Allow the air of the flask to escape out. After an interval of about two minutes put a gas jar filled with water on the bee-hive shelf as shown in Diagram 6.3.

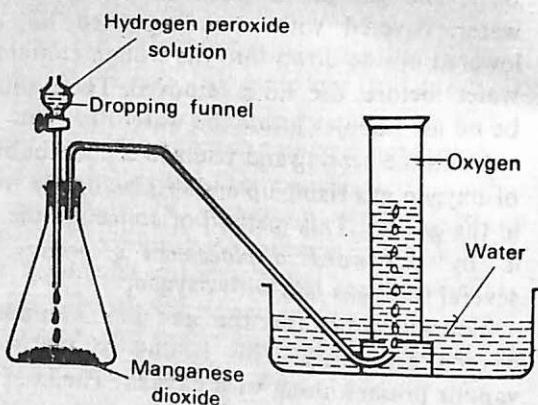


DIAGRAM 6.3 Preparation of oxygen from hydrogen peroxide.

As each drop of hydrogen peroxide solution falls on to the manganese dioxide, a small quantity of oxygen is evolved. Hence the evolution of oxygen can be controlled by regulating the tap of the dropping funnel.

When the gas jar is filled with oxygen by displacing water, cover it with a well-greased glass lid beneath the level of the water in the trough and remove the jar. You can collect some more jars of oxygen in this way.

Since oxygen is collected over water, the gas jars will contain a small amount of water vapour along with oxygen.

6.5 Properties of Oxygen

Physical Properties

1. Oxygen is a colourless and odourless gas.
2. It is slightly soluble in water (about 3%). Living organisms breathe the oxygen dissolved in water.
3. Add two drops of solutions of *blue* and *red litmus* in two separate jars of oxygen. There is no change in the colour of the litmus solutions. Hence the gas is *neutral*.

Chemical Properties

1. COMBINATION WITH NON-METALS

Experiment 6: Place a piece of dry phosphorus (non-metal) in a deflagrating spoon. Warm it and lower it into a jar containing oxygen (Diagram 6.4). The phosphorus burns brightly with a dazzling yellow flame and produces dense white fumes of phosphorus pentoxide. The intensity of burning is greater in oxygen than in air.

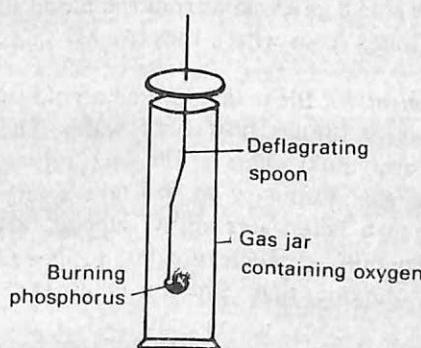
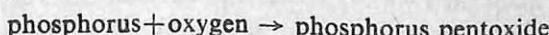


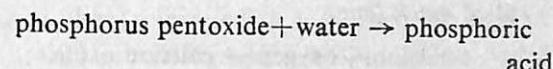
DIAGRAM 6.4 Burning of phosphorus in oxygen.

Oxygen is the gas which supports the burning of substances in air. This property is diluted because of the presence of nitrogen in air. The reaction for the burning of phosphorus is written as follows:

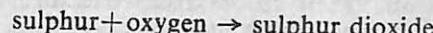


Add a small amount of water in the jar containing white fumes of phosphorus pentoxide and shake well with the jar covered with a greased lid.

The white fumes disappear by dissolving in the water. Add two drops of blue litmus to this solution. The litmus solution turns red. This is the property of an *acidic* solution. Phosphorus pentoxide dissolves in water to form *phosphoric acid* and the reaction may be written as follows:



Experiment 7: Take powdered sulphur in a deflagrating spoon. On heating, it changes into a dark brown liquid and then burns with a blue flame. Introduce the burning sulphur into a jar of oxygen. A colourless and very pungent gas, sulphur dioxide, is formed:



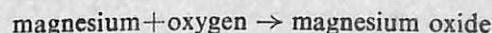
Put a little water into the gas jar and shake well to dissolve the gas. When two drops of blue litmus are added to the solution, it turns red. An acidic solution of sulphurous acid is formed:



Phosphorus and sulphur are non-metallic elements. Their oxides are called *acidic oxides* because their solutions in water are acidic.

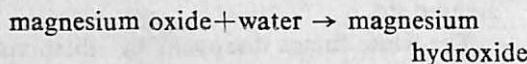
2. COMBINATION WITH METALS

Experiment 8: Hold a piece of magnesium ribbon with a pair of tongs. Burn it in a bunsen flame and immediately introduce the burning ribbon into a jar of oxygen. The ribbon burns with a blinding white light and produces dense white fumes of magnesium oxide:

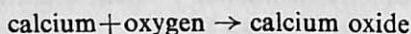


Add some hot water to the jar to dissolve the oxide, which is slightly soluble in water. When two drops of *red litmus* are added to the solution the litmus turns blue. This is the property of

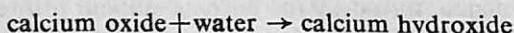
a basic solution. Magnesium oxide forms magnesium hydroxide, which is basic in character, on reacting with water:



Experiment 9: Take a small piece of calcium in a deflagrating spoon. Heat it and lower it into a jar of oxygen. Calcium burns with a brick-red flame forming a white solid, *calcium oxide*, also called *quick lime*:

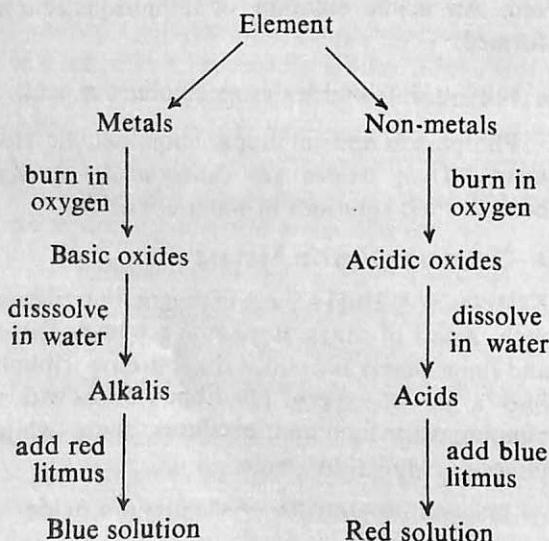


Add a small amount of water to the solid formed and shake well. Calcium oxide dissolves slightly to form calcium hydroxide. The solution of calcium hydroxide turns red litmus blue:



Both magnesium and calcium are metallic elements. Their oxides dissolve to form a basic solution called an *alkali*. Such oxides are called *basic oxides*.

We can summarise the formation of oxides and their solutions in water as follows.



Litmus can thus be used to find whether a solution is acidic, basic or neutral. Such substances which change colour in acidic and basic

solutions are called indicators. We shall learn more about other indicators later on.

6.6 Oxidation and Reduction

You have seen that metals and non-metals combine with oxygen to form their oxides. A reaction in which oxygen combines with another substance is called oxidation. The substance which combines with oxygen is said to be oxidised. Magnesium, for example, is said to be oxidised when it forms magnesium oxide by gaining oxygen. Burning in air is an oxidation reaction.

A reaction in which oxygen is removed from a substance is called reduction. For example, when mercury oxide is heated, it loses oxygen and is said to be reduced by heating.

Oxidation and reduction have wider meanings which we shall study in the coming chapters.

6.7 Respiration and Oxidation

In respiration, air is inhaled into the lungs, from where a portion of it enters the blood. A part of this air reacts with the digested food and the waste gases pass from the blood stream to the lungs from where they are exhaled out.

Experiment 10: Blow the exhaled air through a clean glass tubing into lime water. The lime water turns milky. This is the test for carbon dioxide gas. Ordinary air will take a very long time to turn lime water milky. Hence, exhaled gases contain a much greater proportion of carbon dioxide than present in air (Diagram 6.5).

A small portion (about 4 per cent) of the oxygen we breathe in, is used up to oxidise the food. The slow oxidation of food produces heat energy which keeps our body warm and is used for muscular activities. The energy liberated in the oxidation of food does not, however, produce high temperature and light which are associated with burning. The exhaled gases contain about 17 per cent of oxygen as compared

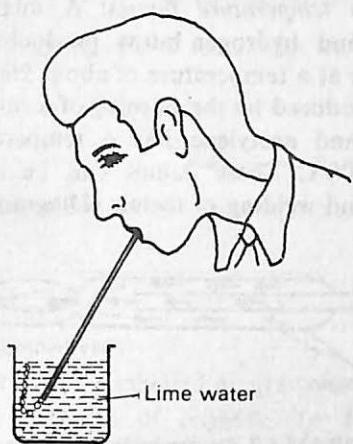


DIAGRAM 6.5 Exhaled air turns lime water milky.

to about 21 per cent present in ordinary air. The composition of inhaled and exhaled air by volume is shown in Table 6.1.

TABLE 6.1 Composition of inhaled and exhaled air by volume.

Composition of inhaled air	Composition of exhaled air
Nitrogen	78.1%
Oxygen	20.9%
Carbon dioxide	0.03%
	78.1%
	17%
	4%

6.8 Large Scale Preparation of Oxygen

Oxygen has vast applications in everyday life. It is required in very large quantities by industries. The cost of production is a very important consideration while preparing a substance on a large scale. The laboratory preparation methods, described earlier, in which certain compounds are heated to obtain oxygen, are not cheap for obtaining large quantities of the gas.

Air is clearly the cheapest and most widely available source of oxygen. Oxygen is obtained from air by first cooling air to very low temperatures so that it condenses to give *liquid air*.

The liquid air is then heated carefully to drive off other gases with higher boiling points. Oxygen, in the liquid form, is left behind.

If you allow compressed air to suddenly come out of an inflated football and if you put your finger over the valve, you will feel a cooling effect. The same principle is applied to change air into the liquid state.

Air, free from carbon dioxide and water vapour, is compressed so that it exerts a pressure of 200 times the normal atmospheric pressure. The compressed air is allowed to expand through a small hole. Intense cooling takes place. The cooled expanded air is used to cool a fresh lot of compressed air, which on expansion will cool further. This is repeated till the air is cooled to such an extent that it changes to liquid air.

The temperature of liquid air is about -200°C . The liquid air is allowed to evaporate slowly when nitrogen, which has a boiling point of -196°C , escapes out as gas leaving behind liquid oxygen, which has a boiling point of -183°C . The liquid air is then filled into strong steel cylinders (Diagram 6.6).

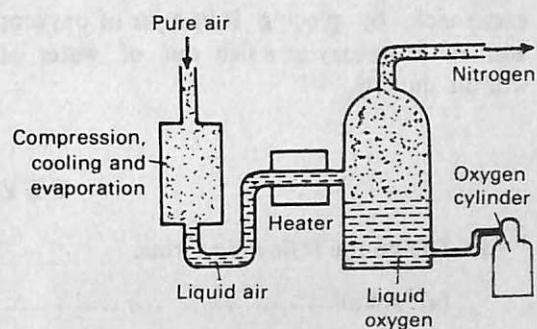


DIAGRAM 6.6 Large scale production of oxygen.

Nitrogen is also obtained in this method. If very pure oxygen and nitrogen are needed, the liquid air is subjected to fractional distillation.

6.9 Uses of Oxygen

The uses of any substance depend upon its pro-

perties. In the study of some of the properties of oxygen, therefore, a reference has already been made to its uses.

(a) *In breathing:* Oxygen present in air is a supporter of life. By oxidising food, it produces heat energy which keeps our body warm and enables us to do work. Patients in hospitals suffering from certain diseases of the lungs are provided increased proportion of oxygen from oxygen cylinders. At high altitudes the air becomes rarefied and the supply of oxygen is not adequate. For persons going on high mountain expeditions and for high altitude flying, an extra supply of oxygen from oxygen cylinders becomes necessary. Fishes breathe in the oxygen dissolved in water while deep-sea divers carry oxygen equipment with them.

It would be well worth remembering that too much supply of oxygen would increase the rate of energy production which may be harmful for a living organism. A fish adjusted to breathing the small quantity of dissolved oxygen from water would soon die when kept in the open air. You can experiment with a small living organism adjusted to normal air, such as a cockroach, by placing it in a jar of oxygen; it will feel as uneasy as a fish out of water and will die quickly.

(b) *High temperature flames:* A mixture of oxygen and hydrogen burns, producing a very hot flame at a temperature of about 2000°C . A flame produced by the burning of a mixture of oxygen and acetylene has a temperature of about 3300°C . These flames can be used for cutting and welding of metals (Diagram 6.7).

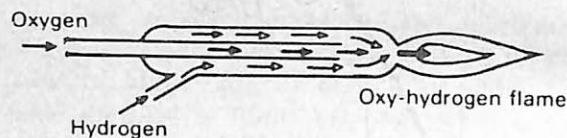


DIAGRAM 6.7 An oxy-hydrogen flame torch.

(c) *Fuel:* Liquid oxygen is used in rockets to make the burning of fuel more vigorous. Liquid oxygen is preferred to gaseous oxygen because it has much less volume (greater density) for the same mass of oxygen.

(d) Oxygen, mixed with a gaseous fuel, produces a flame which is used for giving shape to glass apparatus by melting and moulding the glass.

(e) *Purification of iron:* Cast iron is an impure form of iron; it can be changed into steel by using compressed oxygen.

EXERCISES

1. Define the following terms.

- (a) Alkali.....
.....
.....
- (b) Catalyst.....
.....
.....
- (c) Oxidation and Reduction.....
.....
.....

2. Give differences and similarities between respiration and combustion.

Similarities.....
.....
.....

Differences

.....
.....

3. (a) Draw a labelled diagram for the preparation of oxygen by hydrogen peroxide.

- (b) Why is the gas jar filled with water for collecting the gas, put in the trough after a few minutes?

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.....
.....

- (c) Is the gas collected pure? Give reason
for your answer.....

.....
.....
.....

- (d) Write the word equation for the reaction.....

.....

4. (a) What are indicators?.....

.....

- (b) How would you find whether a solution is acidic, basic or neutral?.....

.....
.....

(c) Why is the preparation of oxygen by hydrogen peroxide in the laboratory a better method than preparing the gas from potassium chlorate?

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.....

(d) Write the word equation for the preparation of oxygen from potassium chlorate

.....

5. (a) Briefly describe the principle of obtaining oxygen on a large scale

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.....

(b) Give two industrial uses of oxygen.....

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.....

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6. (a) What is the difference in the composition of inhaled and exhaled air?.....

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(b) How will you show experimentally that manganese dioxide acts as a catalyst in the preparation of oxygen from hydrogen peroxide?.....

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(c) How will you distinguish between oxygen, chlorine and carbon dioxide gases?
.....
.....
.....

(d) Catalysts increase the rate of a reaction. Is this statement true or false? Give reasons for your answer.
.....
.....

7. Describe what you would observe when the following are heated in a deflagrating spoon in oxygen and the product is treated with water. Write word equations for each.

(a) *Sulphur*

(i) Observations.....
.....
.....

(ii) Equations.....
.....

(b) *Phosphorus*

(i) Observations.....
.....

(ii) Equations.....
.....

(c) *Calcium*

(i) Observations.....
.....

(ii) Equations.....
.....

(d) *Sodium*

(i) Observations.....
.....

(ii) Equations.....
.....

Water and Hydrogen

Water is the most widely distributed liquid on the earth. About 70 per cent of the earth's surface is covered with water of lakes, rivers, seas and oceans. Water is present as vapour in the air and as clouds in the upper atmosphere. The soil contains large quantities of water which is essential for the growth of plants. Without water all human beings and animals would die of thirst. It is essential for cooking and washing.

Our food contains a large quantity of water, from 3 to 4 per cent in dry cereals such as maize and wheat, to about 90 per cent in tomatoes and fresh green vegetables. Milk contains about 88 per cent of water. In fact, our body contains nearly 65 per cent of water. It is impossible to imagine life without water.

7.1 The Water Cycle

The circulation of water from the earth's surface to the atmosphere and back to the earth is called the *water cycle* (Diagram 7.1). Water is constantly going into the atmosphere by evaporation from oceans, rivers and lakes. Plants give out large amount of water vapour during *transpiration*. This water vapour condenses and

returns to the earth in the form of rain, snow and hail.

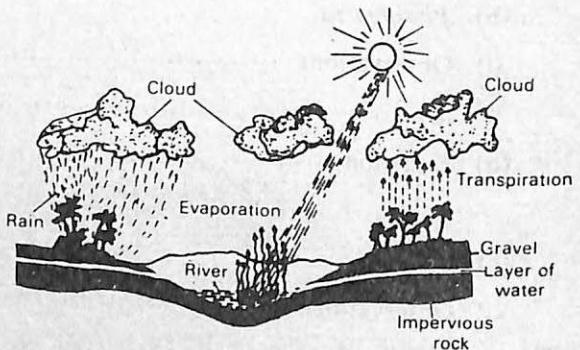


DIAGRAM 7.1 The water cycle.

Air is a great storehouse of water vapour. Attempts are being made by scientists to utilise the water vapour in the air to create artificial rain which will be helpful in times of drought.

7.2 Water in Nature

In nature, water occurs mainly in the following forms.

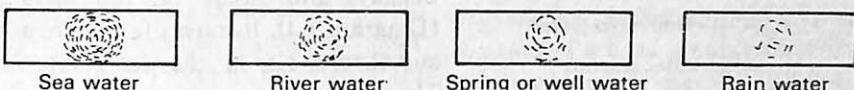
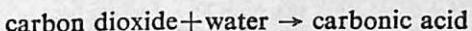


DIAGRAM 7.2 Residue left after evaporation of various types of waters.

1. *Rain water*: Rain water is the purest form of water. The heat of the sun evaporates water from the surface of oceans, lakes and rivers. Since this water vapour is lighter than air, it rises up and condenses to form clouds. It is this water in the clouds which falls down as rain. Thus rain is naturally occurring *distilled water*. However, it dissolves a little of oxygen and nitrogen and a fair amount of carbon dioxide. Carbon dioxide forms carbonic acid with water:



If rain water is evaporated, no residue is left behind except a few dust particles.

2. *Well water and spring water*: When rain falls, part of the rain water sinks into the ground, if the ground is porous, till it reaches a layer of non-porous (impervious) rock. Here it gathers and forms a *reservoir*. If we dig deep enough we can reach this underground water; thus a well is formed.

The rain water may also run along the surface of the impervious rock and come out on the surface elsewhere as *spring water*.

Well water and spring water are both free from suspended impurities because they have been filtered by nature in passing through many layers of sand and rock. They, however, contain dissolved salts. These salts are not poisonous; in fact many of these salts have medicinal values.

3. *River water*: The water on the surface of the earth finally runs into rivers. It contains many dissolved and suspended impurities. There will also be particles of dead plants and animals living on the soil. Several kinds of tiny living bacteria may also be present, some of which cause dangerous diseases such as typhoid,

dysentery and cholera. Hence river water must be sterilised to make it fit for drinking.

4. *Sea water*: Streams and rivers run into the sea. Hence sea water is the most impure form of natural water. It contains all types of impurities present in river water in greater quantities.

Experiment 1: Put two drops each of rain water, spring water or well water, river water, and sea water, if available, on separate microscope slides. Evaporate the samples of water by putting the slides in the sun.

You will observe that rain water leaves hardly any residue whereas sea water leaves the maximum residue (Diagram 7.2). This shows that sea water contains the maximum amount of impurities and rain water the minimum amount.

7.3 Town Water Supply

A town usually gets its water supply from rivers. Before the water is fit for drinking it must be freed from (a) suspended impurities, and (b) bacteria. The following purification methods are employed.

1. *Sedimentation*: Water is allowed to flow through a number of settling tanks where bigger particles settle at the bottom of the tanks by the process called *sedimentation*. Fine particles, such as those of clay can be made to settle by adding alum.

2. *Filtration*: The clear water thus obtained is then passed through a sand filter where any left over suspended impurity is removed. This filter consists of layers of sand, gravel and stones (Diagram 7.3).

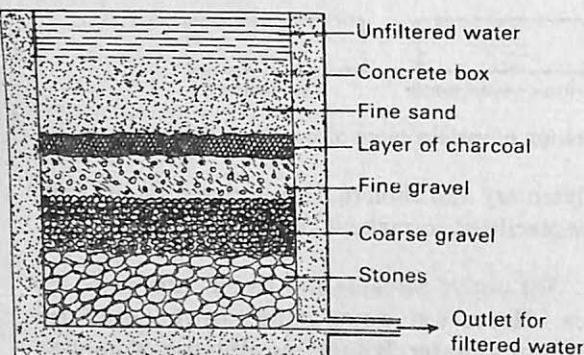


DIAGRAM 7.3 Sand filter.

3. Treatment against bacteria: Bacteria are small organisms which are invisible to the naked eye, and can be seen only through a microscope. Removal of these organisms is known as *sterilisation*. The filtered water is treated with chlorine which destroys the bacteria. Only a small amount of chlorine is needed for this purpose. Swimming pools are chlorinated to remove the dangers of bacterial infection. In some advanced countries ozone is used to kill bacteria. Drinking water for domestic use can be sterilised by boiling.

7.4 Action of Water on Metals

1. Reaction with Sodium

Experiment 2: Sodium metal is always kept in kerosene oil as it reacts violently with water. The water vapour present in air can also react with the metal. Even moisture of the hand can react with sodium. It is, therefore, never touched by hand.

With the help of a pair of tongs take out a piece of sodium from a kerosene container and put it on a filter paper. Cut off a small piece of the metal of about 2 mm side, and dry the piece with the filter paper. Drop the metal into water taken in an evaporating dish. Keep your face away from the dish.

You will observe that the metal is light (low density) and floats on the surface of water (Diagram 7.4). It moves fast all over the surface and there is rise in the temperature of water. Thus, this is an *exothermic reaction* in which heat energy is released. The metal changes into a silvery ball which disappears after some time. There is brisk effervescence showing that a gas is produced. You may also see sparks of yellow colour.

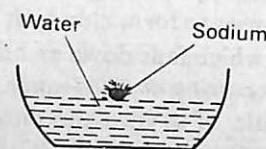


DIAGRAM 7.4 Reaction of sodium with water.

Take in a test tube about 2 cm^3 of the liquid left in the dish and add to it a drop of red litmus. The litmus turns blue showing that an alkali is formed by the action of sodium on water.

Experiment 3: Wrap a small piece of dry sodium in an iron wire gauze. Drop it into a dish containing water and immediately cover it with an inverted test tube full of water. Bubbles of gas will be seen rising up into the test tube. The gas is collected in the test tube by downward displacement of water (Diagram 7.5).

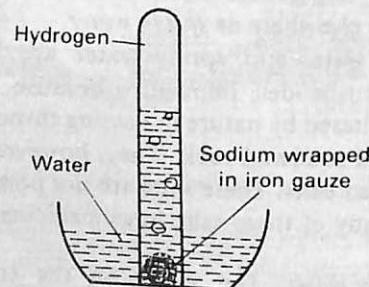
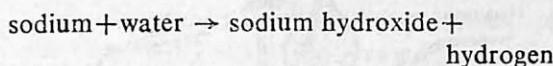


DIAGRAM 7.5 Sodium reacts with water to give hydrogen gas.

Close the test tube with your thumb under the surface of water and remove it from the dish. Introduce a burning splint into the test tube.

The gas burns with a 'pop' sound. This gas is *hydrogen*. It could not have come from sodium which is an element. The gas has come from water which is a compound of hydrogen. The alkali formed is a solution of *sodium hydroxide*. The reaction can be written as follows:



2. Reaction with Calcium

Experiment 4: Drop a piece of calcium, a grey metal, into a dish containing water. Unlike sodium, the metal sinks at the bottom as it is denser than water. You will notice effervescence taking place. Invert a test tube full of water over the metal. Bubbles of gas rise and get collected by displacing water (Diagram 7.6). When all the water in the test tube has been displaced, close the open end of the test tube with your finger and remove it from the dish.

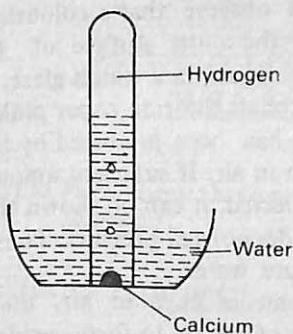
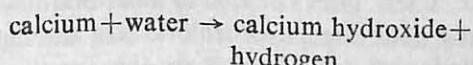


DIAGRAM 7.6 Calcium reacts with water to give hydrogen.

Introduce a burning splint into the test tube. The gas burns with a 'pop' sound, showing that it is hydrogen. A milky solution is left in the dish which turns red litmus blue showing that the solution is basic. The milky solution is a suspension of white particles of calcium hydroxide. The equation can be written as follows:



3. Reaction with Magnesium

Experiment 5: Put about 1 cm³ of water at the bottom of a hard glass test tube and add asbestos wool until the water is soaked up by the asbestos. Fix the test tube on a clamp stand and introduce a small amount of powdered magnesium metal as shown in Diagram 7.7. Heat the metal and the asbestos wool alternately until magnesium starts burning.

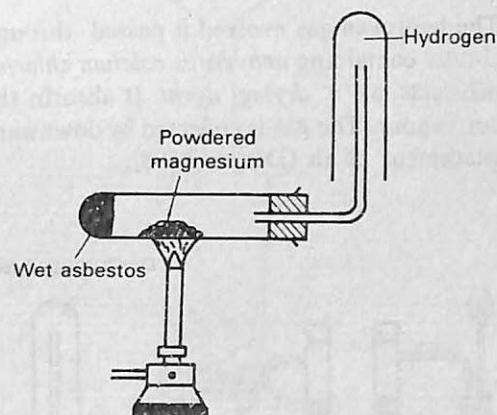
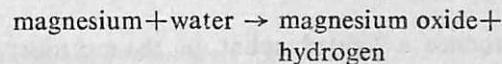


DIAGRAM 7.7 Magnesium reacts with steam to give hydrogen gas.

Collect the gas given out in an inverted test tube and introduce a lighted splint into the test tube. The gas burns with a 'pop' sound, showing the presence of hydrogen gas. A white solid is left in the test tube. This is magnesium oxide.

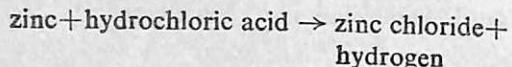


The reactions of water with metals thus show that metals have different reactivities. Sodium reacts very vigorously with cold water, calcium reacts less vigorously with cold water. Heated magnesium reacts only with hot water or steam.

7.5 The Synthesis of Water

Synthesis is the method of preparing a complex substance from its elements or from simpler substances. Let us study how water can be prepared from two elements, hydrogen and oxygen.

Experiment 6: Prepare hydrogen gas by the action of dilute hydrochloric acid on zinc. The acid is poured into a flat flask containing zinc through a thistle funnel until the lower end of the funnel dips in the acid. This is to make sure that hydrogen gas does not escape out through the delivery tube. A solution of zinc chloride is left in the flask.



The hydrogen gas evolved is passed through a U-tube containing *anhydrous calcium chloride* which acts as a *drying agent*. It absorbs the water vapour. The gas is collected by downward displacement of air (Diagram 7.8).

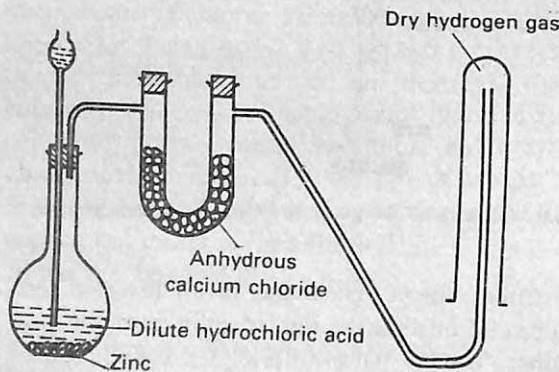


DIAGRAM 7.8 Preparation of hydrogen.

Introduce a lighted splint in the test tube. The gas is seen to burn with a loud sound. This shows that there is some air mixed with the hydrogen gas, as air and hydrogen form an explosive mixture. Continue to take fresh samples of hydrogen gas in the test tube and light the gas until it burns quietly. This indicates that air has been removed from the apparatus.

Light the gas at the mouth of a jet and allow the flame to play on the cold surface of a glass retort through which a stream of cold water flows continuously (Diagram 7.9).

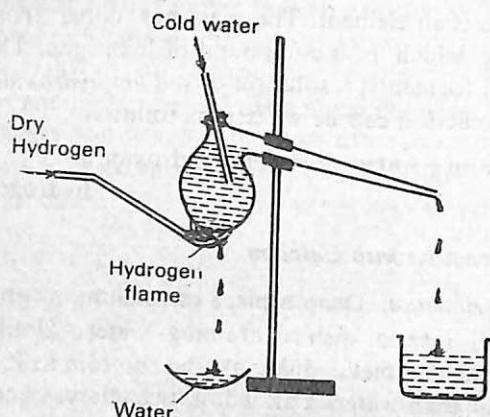
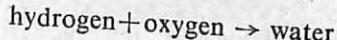


DIAGRAM 7.9 Hydrogen burns in air to give water.

Remember that hydrogen should not be lighted at the jet until all the air has been displaced from the flask. Otherwise a serious explosion of the apparatus may take place.

You will observe that a colourless liquid is formed on the outer surface of the retort. Collect the liquid in a watch glass. The liquid turns blue cobalt chloride paper pink, showing that water has been produced by the burning of hydrogen in air. If sufficient amount of this liquid is collected, it can be shown that it boils at 100°C under normal pressure. Thus the liquid formed is pure water.

When elements burn in air, they combine with oxygen of the air to form oxides. Hydrogen, therefore, combines with oxygen to form an oxide of hydrogen which is water:



7.6 Preparation of Hydrogen in the Laboratory

Experiment 7: A convenient method for preparation of hydrogen gas in the laboratory is by the action of dilute hydrochloric acid on zinc.

Take a few pieces of zinc in a flat-bottomed flask and pour dilute hydrochloric acid through a thistle funnel until the end of the funnel is

dipping in the acid. This is to ensure that hydrogen does not escape through the funnel.

A brisk effervescence indicates the formation of hydrogen gas. Allow some gas to escape out because it is mixed with air inside the flask. Collect the gas by downward displacement of water (Diagram 7.10). Put a greased glass plate over the open end of the gas jar and remove it from the trough. The gas collected still contains water vapour as an impurity. By placing other gas jars filled with water on the bee-hive shelf, several jars containing hydrogen can be collected.

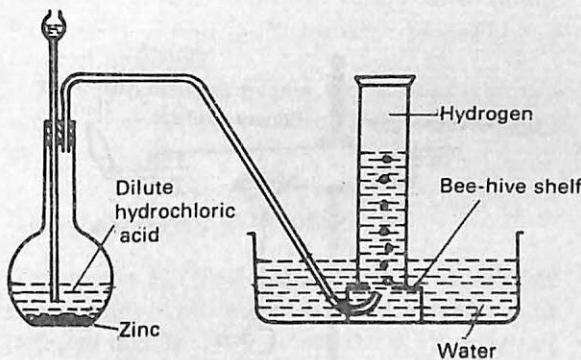


DIAGRAM 7.10 Preparation of hydrogen in the laboratory.

7.7 Properties of Hydrogen

Physical Properties

1. Examine a gas jar of hydrogen. The gas is colourless and odourless.
2. Introduce pieces of blue and red litmus papers into a jar of hydrogen. You will observe no change of colour showing that the gas is neutral.
3. Fill a small amount of dilute hydrochloric acid in a bottle and drop in few pieces of zinc. After a while, fit a balloon on the mouth of the bottle (Diagram 7.11). When the balloon is inflated, tie it with a string. If you release the balloon, it rises up showing that the gas is lighter than air. In fact, it is the lightest gas known.

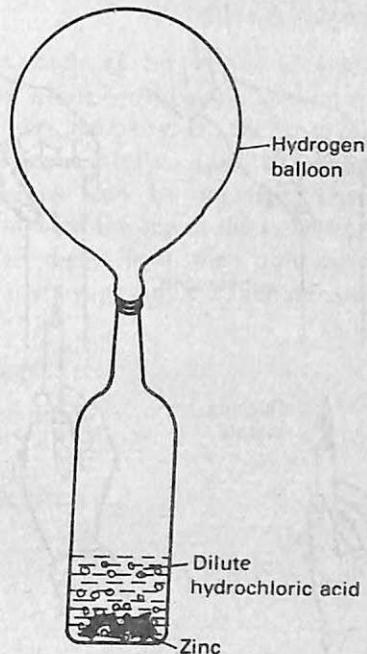
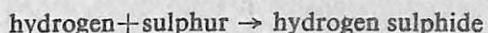


DIAGRAM 7.11 Filling a balloon with hydrogen gas.

4. Hold a jar filled with hydrogen with its mouth downwards. Hold a lighted candle and put the flame inside the jar. The candle is extinguished and the gas burns with a 'pop' sound. Hydrogen is a flammable gas but does not support combustion (Diagram 7.12).

Chemical Properties

1. *Experiment 8—combination with sulphur:* Bubble hydrogen gas through molten sulphur taken in a porcelain dish. A gas is evolved which smells like rotten eggs. The gas is hydrogen sulphide:



2. *Experiment 9—combination with chlorine:* Place a gas jar of chlorine, a greenish-yellow gas, over a jar of hydrogen, mouth to mouth. Allow the gases to mix and expose the gas jars to diffused sunlight (Diagram 7.13).

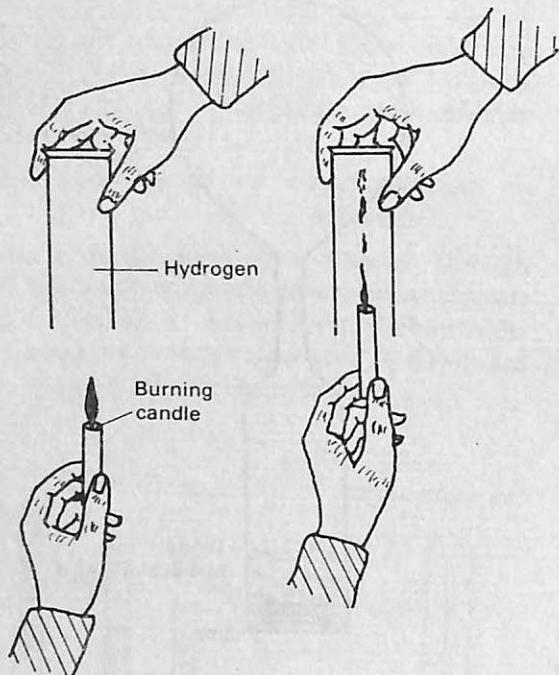


DIAGRAM 7.12 Hydrogen does not support combustion.

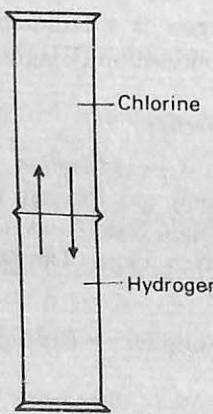
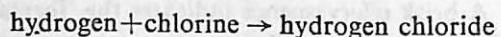


DIAGRAM 7.13 Reaction of hydrogen and chlorine.

The greenish-yellow colour of chlorine disappears and a colourless gas is formed in both

the gas jars. The gas is hydrogen chloride and it fumes in air:



3. Experiment 10—action on metallic oxides:

Pass a stream of dry hydrogen over black copper oxide contained in a hard glass tube. Collect samples of hydrogen at intervals in test tubes and introduce a lighted splint. When the gas burns quietly, it shows that the air has been displaced from the apparatus. Now burn the hydrogen at the jet on the other end of the tube. Heat the copper oxide from below (Diagram 7.14).

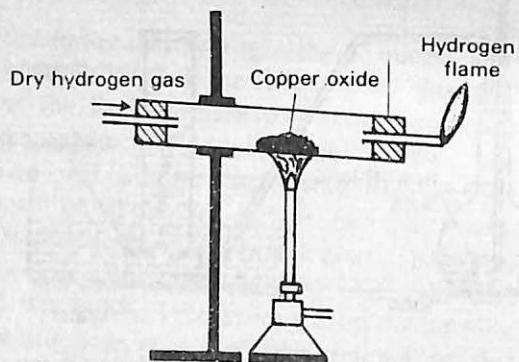
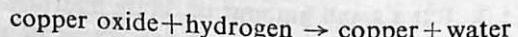


DIAGRAM 7.14 Hydrogen reacts with copper oxide to form copper.

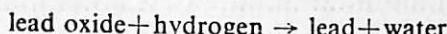
Remove the burner after about 20 minutes of heating, but continue to pass the hydrogen gas while the solid inside the tube is cooling. The residue left is reddish brown in colour and is the metal copper. (The hard glass tube is kept slightly tilted to prevent the water formed in the reduction from running back to the hot part of the tube.)



As the copper is cooling, the stream of hydrogen is passed over the metal. This is to avoid air entering the tube and oxidising the hot copper metal.

A similar result is obtained when hydrogen gas is passed over heated lead oxide, an orange

yellow solid. The products formed are grey metallic lead and water:



7.8 Uses of Hydrogen

1. One of the important uses of hydrogen is that in combination with nitrogen it forms *ammonia*, which forms the basis of many fertilisers. Fertilisers are extensively used to get a better yield from crops.

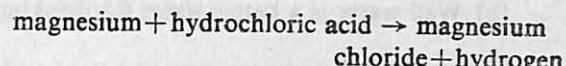
2. Hydrogen gas is used to convert liquid oils, such as groundnut and coconut oils, into solid fats. This process is called *hydrogenation of oils* and takes place in the presence of nickel as a catalyst at 200°C .

3. Oxy-hydrogen flames which have a temperature of about 2000°C are used for welding of metals.

7.9 Reactivity of Metals

Experiment 11: You have noticed in earlier experiments that water reacts differently with different metals. Take about 0.5 g of each of four metals: magnesium ribbon, powdered zinc, iron dust and copper turnings, in four separate test tubes. Add 5 cm^3 of dilute hydrochloric acid to each of the metals and note the intensity of the reaction.

You will notice that magnesium reacts most vigorously and disappears first. The reaction of zinc is less vigorous while iron shows even less reactivity. Copper does not react at all with dilute hydrochloric acid. The reaction of the acid with magnesium, zinc and iron produces hydrogen:



Reactions such as the action of water and the action of dilute hydrochloric acid on metals differ in their reactivity. On the basis of such reactions, a list of metals, called the *reactivity series of metals*, can be prepared. The most reactive metal is at the top in the series and the least reactive metal is at the bottom. This arrangement showing some of the metals is as follows.

1. Potassium
2. Sodium
3. Calcium
4. Magnesium
5. Aluminium
6. Zinc
7. Iron
8. Tin
9. Lead
10. [Hydrogen]
11. Copper
12. Mercury
13. Silver
14. Gold
15. Platinum

Hydrogen, although a non-metal, has been included in the series because in some ways it shows metallic properties, e.g. like metals, it forms positively charged particles.

The more reactive metals cannot occur in the free state in nature. Gold can occur in free state in nature and retains its shine because of its resistance to combine with other elements.

EXERCISES

1. Explain the following, giving reasons.

- (a) Rain water is as good as distilled water.....

(b) Well water is a better water for drinking purposes than river water.....

.....
.....
.....

(c) Sodium metal should not be touched by hand

.....
.....

(d) Sodium should not be used in the preparation of hydrogen by the action of dilute hydrochloric acid on a metal.....

.....
.....
.....

(e) When calcium is put into water, the latter becomes turbid.....

.....

2. Which of these statements are true and which are false. Write T for true statements and F for false statements.

- (a) Sodium metal is kept in an ordinary bottle like zinc.
- (b) Magnesium oxide cannot be reduced by hydrogen.
- (c) Metals are never found in free state in nature.
- (d) Copper reacts with hydrochloric acid to produce copper chloride and hydrogen.
- (e) Hydrogen and air form an explosive mixture when burnt together.

3. Define or explain the terms given below.

(a) Exothermic reaction

.....

(b) Water cycle.....

.....

(c) Reducing agent

.....

(d) Sterilisation of water.....

.....

4. (a) List the steps involved in purifying water for drinking purposes.....

.....
.....

(b) Arrange the following metals in decreasing order of reactivity: zinc, sodium, copper, iron and calcium.

.....
.....
.....
.....

5. Complete the following statements by filling in the blanks.

(a) Chlorine is added to water for the present in water.

(b) Anhydrous calcium chloride is used as a

(c) The presence of water can be shown by a paper which turns from to

(d) Hydrogen and chlorine combine in presence of

6. Write word equations for the following reactions.

(a) Action of water on sodium.....

.....

(b) Action of water on magnesium

.....

(c) Action of hydrochloric acid on magnesium.....

.....

(d) Action of hydrogen on molten sulphur.....

.....

(e) Burning of hydrogen in air

(f) Action of hydrogen on hot iron oxide.....
.....

7. (a) Describe all that you would observe when sodium metal is put into water contained in a dish.....
.....
.....

(b) How will you show experimentally that hydrogen acts as a reducing agent?.....
.....
.....

(c) Give two uses of hydrogen

.....
.....

8. (a) Draw a labelled diagram for the laboratory preparation of dry hydrogen.

(b) What do you call the method of collecting the gas?

(c) How would you test if a given sample of hydrogen collected in a test tube is not mixed with air?

.....
.....
.....
.....

Acidic and Alkaline Solutions and Salts

8.1 Acidic Solutions

You may recall that non-metals combine with oxygen to form oxides. When these oxides are dissolved in water, the solutions turn blue litmus red. Such oxides are called *acidic oxides* and the solutions are called *acidic solutions*. You will notice that some reagent bottles in the chemistry laboratory are labelled 'acids'. The acids commonly available are hydrochloric acid, sulphuric acid, nitric acid and acetic acid.

8.2 Concentrated and Dilute Acids

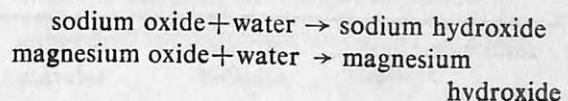
Acids are available either in the concentrated or in the dilute form. Concentrated acid is pure acid containing a little water. Dilute acid contains more water than the acid. While preparing a dilute acid, the pure acid is slowly added to water. Water should never be added to pure acids because sometimes a lot of heat is produced which may crack the glass container.

You should handle concentrated acids with great care; they may produce blisters on the skin and burn holes in clothes. If some concen-

If concentrated acid falls on your body, it should be washed away with plenty of running water.

8.3 Basic Solutions

You may be remembering that metallic oxides such as oxides of sodium and magnesium, when dissolved in water, turn red litmus blue. These oxides are called *basic oxides* and their solutions are called *basic solutions*. The basic oxides when dissolved in water form hydroxides. For example:



Metallic oxides and hydroxides are called bases. A soluble base is called an alkali. The alkalis commonly used in the laboratory are:

1. Potassium hydroxide or caustic potash.
 2. Sodium hydroxide or caustic soda.
 3. Calcium hydroxide (slightly soluble in water, the solution being known as lime water).
 4. Ammonium hydroxide or ammonia solution.

Solutions which are neither acidic nor basic are called *neutral solutions*.

8.4 Indicators

You have been using litmus to find out if a solution is acidic or basic. Litmus is made by oxidation of plants called lichens. Its colour is either blue or red. A paper soaked in litmus solution and then allowed to dry is called litmus paper. Substances such as litmus which change colour when treated with acidic or basic solutions are called indicators. Many other substances, besides litmus, are used as indicators.

Experiment 1: Take 2 cm³ each of dilute hydrochloric acid and sodium hydroxide solution in two separate test tubes. Add one drop of *methyl orange* solution to each of the two test tubes. You will notice that methyl orange turns red in dilute hydrochloric acid and in sodium hydroxide solution its colour changes to yellow.

Another indicator commonly used is *phenolphthalein*. In an acidic solution it remains colourless and in an alkaline solution it turns pink.

The colour in acidic and alkaline mediums of the three commonly used indicators are summarised in Table 8.1.

TABLE 8.1 Colour of common indicators in acidic and alkaline solutions.

Indicator	Acidic solution	Alkaline solution
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink

Experiment 2: Indicators are complex substances. Some indicators can be prepared from coloured substances available in our surroundings.

Mix 10 cm³ of alcohol to the same volume of distilled water. Grind a few petals of a brightly coloured flower with the aqueous alco-

hol. Heat the contents gently in a conical flask. Decant or filter the coloured extract into a test tube. Add to this two drops of dilute hydrochloric acid. See if there is any change in colour. To the same test tube add a few drops of sodium hydroxide solution and see if a different colour is obtained.

You will notice that some of these coloured extracts give different colours in acidic and alkaline solutions and can be used as indicators. Purple bougainvillea gives good results.

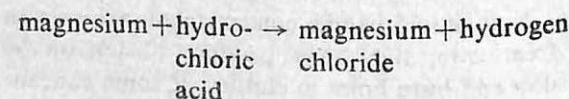
8.5 Characteristics of Acidic Solutions

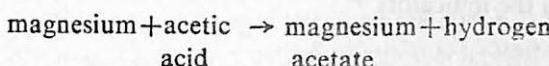
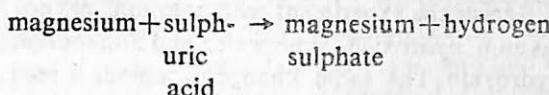
(1) *Experiment 3:* Prepare a very dilute solution of hydrochloric acid and put some drops of this solution on the tip of your tongue.

You will find that the solution has a sour taste. All acids have a sour taste. The sour taste of fruits is due to the presence of some acids. Lemon and citrus fruits contain citric acid, grapes contain tartaric acid while vinegar contains acetic acid. You can cut open a lemon and test the presence of an acid with the help of a blue litmus paper.

(2) *Experiment 4:* Take about 5 cm³ each of dilute hydrochloric acid, dilute sulphuric acid and dilute acetic acid in different test tubes, and add a piece of magnesium ribbon to each of the test tubes. There is effervescence in each test tube. The gas produced is hydrogen because it burns with a 'pop' sound when a lighted splint is introduced into the test tube.

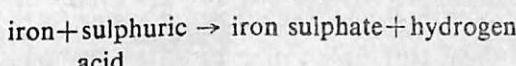
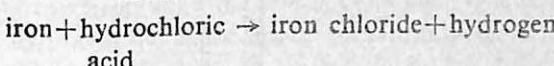
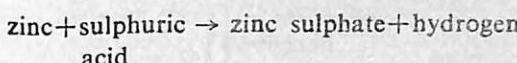
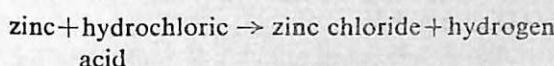
Put one drop each of the solutions left in the test tubes on different microscope slides. Evaporate the liquids by holding the slides over a warm wire gauze. Crystalline solids are left on the microscope slides. These solid residues are called salts of the acids. The reactions may be written as below:





The salts of hydrochloric acid are called *chlorides*, those of sulphuric acid are called *sulphates*. Salts of acetic acid are called *acetates*. Similarly the salts of nitric acid are called *nitrates*, of carbonic acid *carbonates*, and of phosphoric acid *phosphates*.

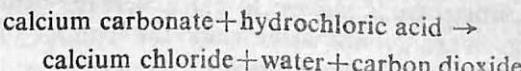
Dilute sulphuric acid and dilute hydrochloric acid react with most of the metals to produce a salt of the acid and hydrogen. Some of the word equations may be written as below:



The action of nitric acid on metals is, however, different.

Thus dilute acids may produce hydrogen on reacting with a more reactive metal.

(3) *Experiment 5:* Take a small amount of solid calcium carbonate in a test tube, and add about 2cm^3 of dilute hydrochloric acid to the calcium carbonate. There will be effervescence and the gas when passed into lime water will turn the lime water milky (Diagram 8.1). The gas evolved is therefore carbon dioxide. The resultant solution on evaporation leaves a white salt, calcium chloride. Water is also formed in the reaction.



Other acids on reacting with carbonates show a similar reaction. The products are salt of the

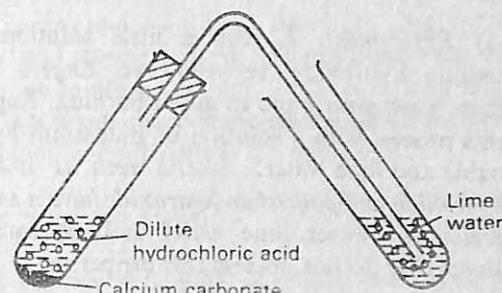
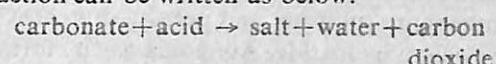
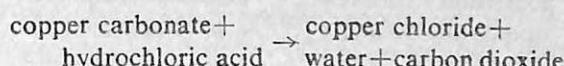
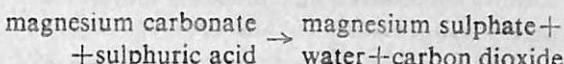
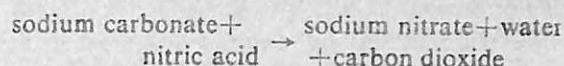


DIAGRAM 8.1 Action of hydrochloric acid on calcium carbonate.

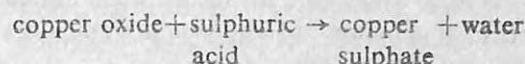
acid, water and carbon dioxide. This general reaction can be written as below:



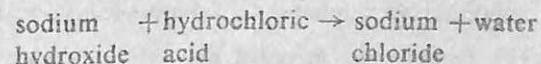
Some examples of such a reaction are as follows:



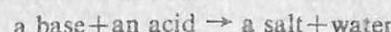
(4) *Experiment 6:* Take some black copper oxide, add dilute sulphuric acid and heat. A blue solution of copper sulphate is formed. Water is the other product of this reaction.



Similarly sodium hydroxide reacts with hydrochloric acid to give a salt, sodium chloride, and water.



A base which is metallic oxide or hydroxide, thus reacts with an acid to form a salt of the acid and water.

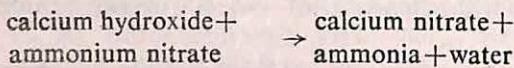
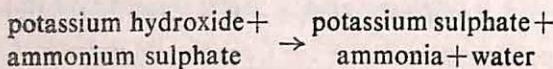
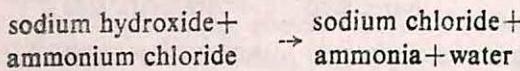


8.6 Characteristics of Alkalies

(1) *Experiment 7:* Rub a little solution of sodium hydroxide between two fingers and then wash your hand to avoid burning. Repeat this process with a solution of potassium hydroxide and lime water. *Alkalies such as sodium hydroxide and potassium hydroxide have a soapy feeling.* However, lime water and ammonium hydroxide do not possess this property.

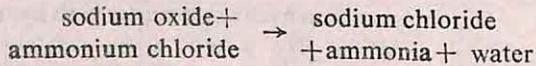
(2) *Experiment 8:* Add sodium hydroxide solution to solid ammonium chloride and heat. A gas is evolved which has a characteristic pungent smell and which turns red litmus blue. The gas is ammonia. Repeat the experiment using ammonium sulphate and potassium hydroxide or calcium hydroxide. Ammonia gas is formed as before.

When alkalis are heated with ammonium salts, ammonia gas is evolved. The reactions can be written as below:



The experiment cannot be performed with ammonium hydroxide which itself gives ammonia on heating.

The basic oxides of these metals when heated with an ammonium salt also produce ammonia gas.



(3) *Experiment 9:* Add the three indicators: methyl orange, red litmus and phenolphthalein to a dilute solution of sodium hydroxide. It will be observed that methyl orange turns yellow, red litmus turns blue and phenolphthalein turns pink.

Repeat the experiment with solutions of potassium hydroxide, lime water and ammonium hydroxide. The same change of colour is seen in the indicators.

8.7 Reaction of an Alkali with an Acid Solution

Experiment 10: Fill a burette with dilute hydrochloric acid and fix it on a clamp stand. Suck in 25 cm³ of dilute sodium hydroxide solution in a pipette and transfer the alkali into a clean beaker. Note the temperature of the sodium hydroxide solution and add a drop or two of methyl orange indicator. Run in the acid solution slowly into the beaker and stir continuously (Diagram 8.2).

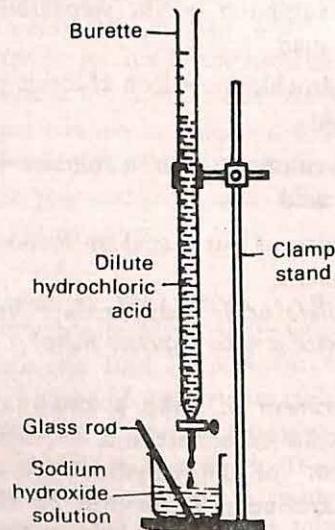
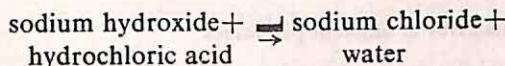


DIAGRAM 8.2 Reaction of dilute hydrochloric acid with sodium hydroxide.

A stage is reached when an additional drop of the acid solution turns the colour of the indicator from yellow to pink. Note the readings in the burette which gives the volume of the acid needed to just react completely with the alkali. Take the temperature of the resultant solution.

You will find that there is a temperature rise showing that a chemical reaction takes place and is exothermic. The resultant solution is neutral. If this solution is evaporated to dryness a white solid residue is left. This residue is sodium chloride, also called common salt. Water is another product of this chemical reaction.



The reaction of an acid with a base to form a salt and water is called neutralisation.

8.8 Strength of Acidic and Alkaline Solutions

Acids and alkalis can be strong or weak. These words have a different meaning than concentrated and dilute. We should be careful in differentiating between these two concepts.

Experiment 11: Take about 2 cm³ each of dilute hydrochloric acid, dilute sulphuric acid and dilute acetic acid in three different test tubes. To each of them add the same length of magnesium ribbon. You will notice that hydrochloric acid reacts more vigorously than sulphuric acid, while the reaction of acetic acid is least vigorous.

Now take about 2 cm³ each of dilute hydrochloric acid and concentrated acetic acid in different test tubes and add the same length of magnesium ribbon to each of the acid solutions. It is observed that dilute hydrochloric acid reacts more vigorously than concentrated acetic acid. Hence hydrochloric acid is stronger than sulphuric acid and sulphuric acid is more strong than acetic acid.

The strength of an acid depends upon its reactivity, while the concentration of the acid depends on the quantity of water in acid. An acid may be more concentrated (containing less water) but may be weaker (less reactive) than a dilute solution of another strong (more reactive) acid. This is very evident from the fact that dilute hydrochloric acid contains more water but

is stronger than concentrated acetic acid which contains less water.

Like in acid solutions, the word concentration and strength have different meanings for alkaline solutions as well. Potassium hydroxide and sodium hydroxide are strong alkalis whereas ammonium hydroxide is a weak alkali.

8.9 The pH Scale

The relative strengths of acidic and alkaline solutions are measured on a scale known as *pH scale*. The pH of an acid or alkali can be found out with the help of pH paper or a universal indicator.

Experiment 12: Put a drop each of different acidic solutions, distilled water, sodium hydroxide solution and ammonium hydroxide on different strips of pH paper. You will notice that different shades of colour appear on the strips of pH paper. By comparing the colours with the colours given by the manufacturers, you can find the pH of the different solutions.

You will notice that acidic solutions have a pH value less than 7, while alkaline solutions have a pH value greater than 7. A solution which has a pH of 7 is said to be neutral. Distilled water will give a pH of 7.

Hydrochloric acid has lower pH than acetic acid and sodium hydroxide has a higher pH than ammonium hydroxide.

Hence, the lower the pH value of the solution, the stronger is the acid. For alkaline solutions the greater the pH value the stronger is the alkali (Diagram 8.3).

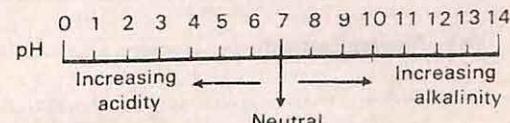


DIAGRAM 8.3 The pH scale.

A *universal indicator* is a mixture of several indicators and is used to find out the strengths

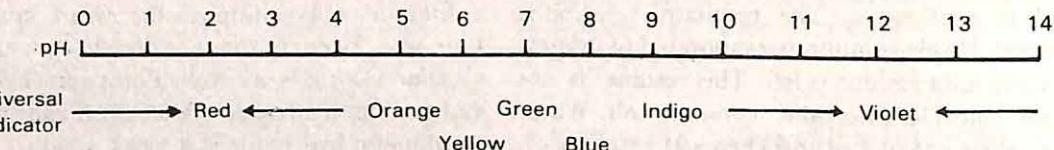


DIAGRAM 8.4 Colours of a universal indicator at different pH values.

of acidic and alkaline solutions. It shows different characteristic colours for each pH value of the solution. One universal indicator shows a green colour in a neutral solution. The colour changes from blue to violet when the pH increases from 7 to 14. In acidic solution the colour changes from yellow to orange and then to red as the pH of the solution decreases (Diagram 8.4).

Experiment 13: Take 1 cm³ each of dilute hydrochloric acid, dilute sulphuric acid, dilute acetic acid, ammonium hydroxide solution and a solution of sodium hydroxide, in different test tubes and dilute each of them with 15 cm³ of distilled water. To each of these test tubes, add about 6 to 8 drops of diluted universal indicator. Shake the test tubes and compare the colours of the solutions with those of the colours given

by the manufacturer. The different colours give the pH of the solutions.

The common indicators such as litmus, methyl orange and phenolphthalein change colour only at certain pH values. Hence these cannot be used to find out the pH of any solution.

Experiment 14: The pH value of soil can be found as follows. Take a sample of the soil in a beaker, add a little distilled water, stir and filter. Test the filtrate with the universal indicator or pH paper to determine the pH of the soil.

The pH value of the soil is of great importance to farmers. There is an optimum (most suitable) pH of the soil needed for a particular crop. For example, citrus fruits grow better in a slightly alkaline soil.

Physicians also make use of pH values of urine and blood for the diagnosis of diseases.

EXERCISES

1. Define or explain the following terms.

(a) Indicator.....

.....

(b) Neutralisation.....

.....

(c) Strength and concentration of acids.....

.....

(d) pH.....
.....
.....

2. Complete the following word equations.

- (a) Sodium hydroxide + nitric acid \rightarrow
- (b) Magnesium carbonate + hydrochloric acid \rightarrow
- (c) Calcium oxide + acetic acid \rightarrow
- (d) Magnesium hydroxide + ammonium sulphate \rightarrow
- (e) Calcium + hydrochloric acid \rightarrow

3. What change of colour would you observe when the following indicators are added to the substances listed below.

<i>Substance</i>	<i>Methyl orange</i>	<i>Litmus solution</i>	<i>Phenolphthalein</i>
1. Dilute hydrochloric acid
2. Sodium hydroxide solution
3. Sodium chloride solution
4. Water

4. Describe all that you would observe when the following experiments are conducted. Write word equations for each reaction.

- (a) Dilute sulphuric acid is added to copper carbonate.....
.....
.....
.....

- (b) Magnesium ribbon is added to acetic acid.....
.....

(c) Potassium hydroxide and ammonium chloride are heated together

.....

.....

.....

5. (a) Describe briefly how you can prepare a sample of pure solid sodium chloride using sodium hydroxide solution and dilute hydrochloric acid.....

.....

.....

.....

.....

(b) Dilute hydrochloric acid is stronger than concentrated acetic acid but is less concentrated. Explain why this is so.....

.....

.....

.....

.....

6. (a) Classify the solutions with the following pH as weak and strong acids or alkalis and neutral solutions.

pH = 3

pH = 12

pH = 7

pH = 8

pH = 6

pH = 5

(b) Which of the following is likely to be water, carbonic acid, hydrochloric acid, ammonium hydroxide, sodium hydroxide.

pH = 3

pH = 12

pH = 7

pH = 8

pH = 6

- (c) Why is it useful to know the pH of a solution?.....
.....
.....
7. (a) How would you find out by a chemical test whether a solution is of an acid or an alkali? (Do not give indicator test.).....
.....
.....
.....
.....
- (b) Why do lemon and vinegar have a sour taste?
-
- (c) Name the products formed when an acid reacts with a base.....
.....
- (d) What precaution should you take when a strong acid falls on your hand.....
.....
- (e) Is there any heat change when an acid reacts with a base? What do you call such reactions?.....
.....

Water and Compounds

You have earlier studied that salts can be prepared by the action of acids on metals, metallic oxides and hydroxides, and metallic carbonates. The solutions containing salts when evaporated produce crystals which have a definite geometrical shape, with smooth faces and sharp edges.

9.1 Water of Crystallisation

Experiment 1: Examine a few crystals of copper sulphate and feel them with your fingers. You will notice that the solid particles have sharp edges. The crystals are absolutely dry.

Take a few crystals in a test tube. Hold the tube with a test tube holder and heat it with its mouth slightly tilted downwards (Diagram 9.1).

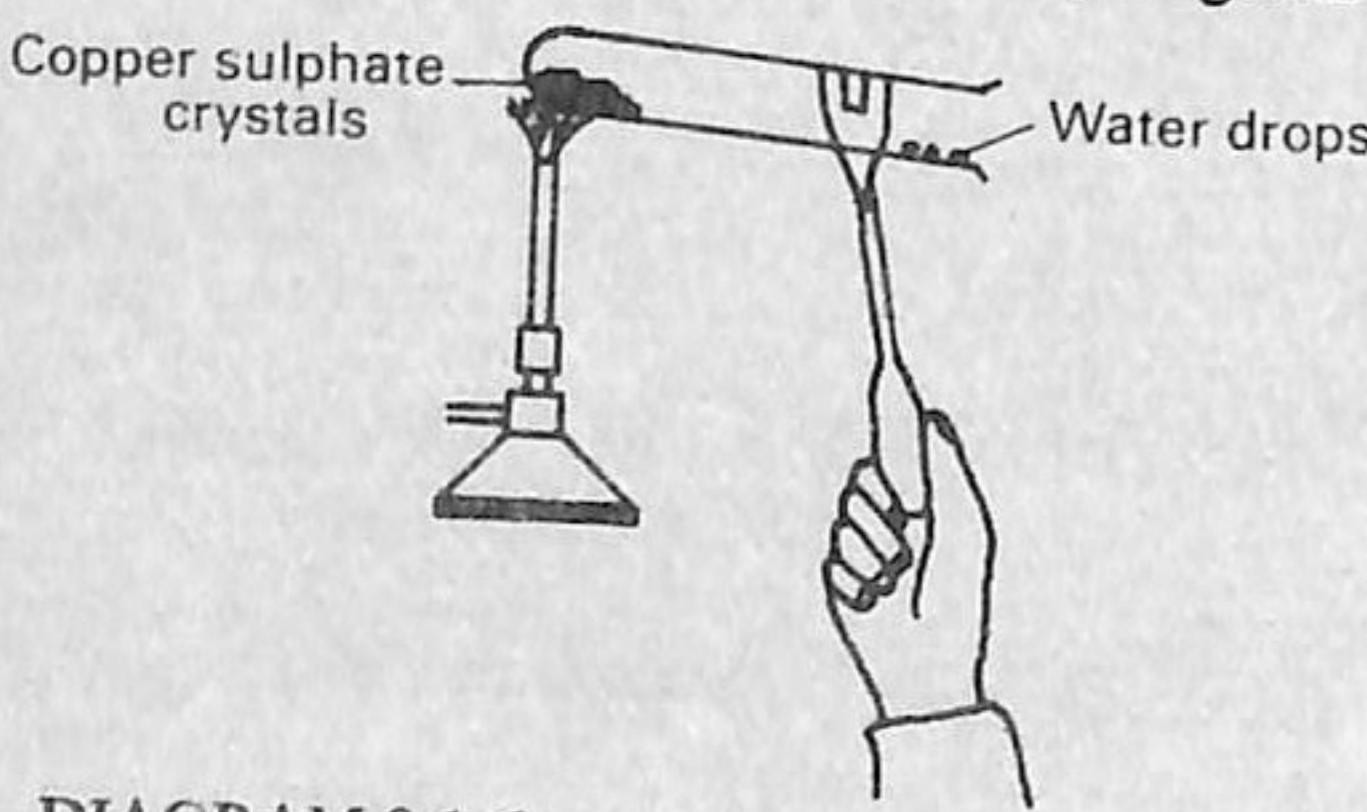


DIAGRAM 9.1 Copper sulphate crystals on heating give water.

You will observe that the blue crystals give a colourless liquid on heating which condenses on the cooler part of the test tube. The liquid turns dry cobalt chloride paper from blue to pink, showing that it is water.

The open end of the tube is inclined slightly downwards to prevent the water from running back to the hot part of the test tube which may be broken.

The blue crystals turn into a white solid. Feel the solid residue with your fingers. You will observe that the solid is powdery and the crystalline structure has been lost. Allow a drop of water to fall on to the white powder and you will notice that the powder turns blue again. A little heat is produced in the process, indicating that an exothermic reaction takes place between the white residue and the water. On drying, blue crystals are again formed.

Thus the crystalline structure is due to the water present in copper sulphate, which is called *water of crystallisation*. It is in loose chemical combination with the compound and is driven out on heating, leaving a white powdery solid residue. Repeat the experiment by heating crystals of washing soda, ferrous sulphate and zinc sulphate

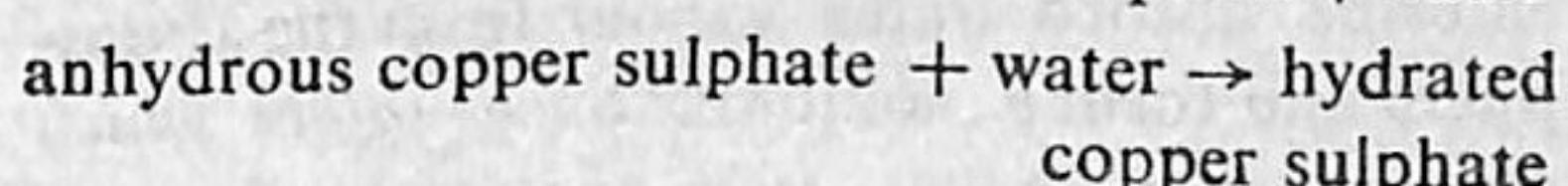
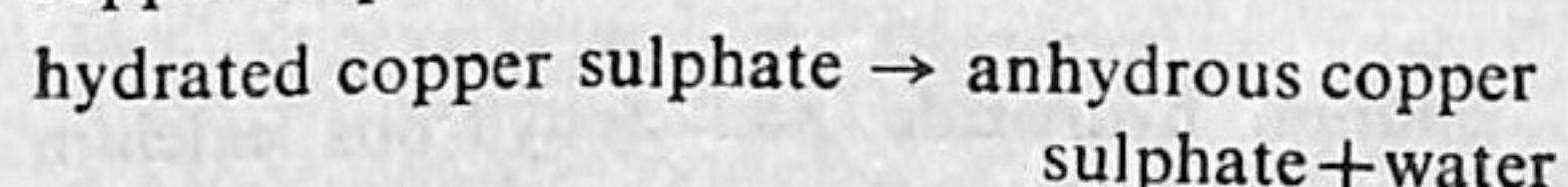
separately. You will observe that these compounds also contain water of crystallisation. Compounds containing water of crystallisation are called *hydrated compounds*. When water is driven out by heating, the solid residue left is called *anhydrous* (without water) and is *amorphous* (non-crystalline) in structure.

Experiment 2: Heat gently a few crystals of pure sodium chloride (common salt) and of potassium nitrate separately in different test tubes. You will observe that crystals of sodium chloride produce a crackling sound called *decomposition* on heating. This is due to bigger particles breaking up into smaller ones. The crystals of potassium nitrate melt into a colourless liquid and on strong heating produce oxygen which rekindles a glowing splint.

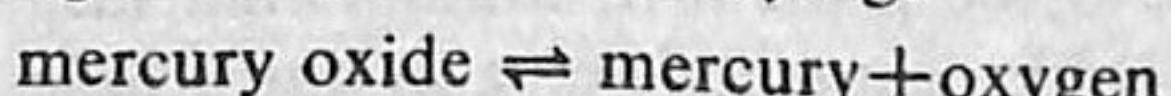
Both these substances, however, give out no water. There are many other compounds which when crystallised from their solutions in water do not contain water of crystallisation, e.g. lead nitrate, ammonium chloride, etc.

9.2 Reversible Reactions

When hydrated copper sulphate is heated, the products are anhydrous copper sulphate and water. On the other hand the reaction between anhydrous copper sulphate and water forms copper sulphate.



Reactions which can go both ways, depending on the conditions, are called reversible reactions. Another example of a reversible reaction is the action of heat on mercury oxide. A reversible reaction is indicated by using double arrows pointing in reverse directions, e.g.



9.3 Saturated Solution and Solubility

Add some powdered potassium nitrate into

some distilled water while constantly stirring the water. You will notice that there is a limit to the amount of the salt that can dissolve in a given amount of water (solvent). Any further amount of potassium nitrate does not dissolve. This solution is called a *saturated solution* at room temperature.

Thus a solution is said to be a saturated solution at a certain temperature when no more solute dissolves in a given amount of the solvent at that temperature in presence of excess of the solute.

If you heat a saturated solution containing a small amount of excess solute, the solute dissolves. On cooling to the original temperature the solute reappears.

The solubility of solid substances generally increases with rise in temperature. In case of a gas its solubility in a liquid depends upon temperature, pressure of the gas and the nature of the gas and the liquid in which it dissolves. The greater the pressure of the gas the greater is the solubility. A soda water bottle when opened gives a lot of effervescence due to bubbling out of carbon dioxide gas. This is because, before opening, the pressure inside the soda bottle is much more than atmospheric pressure and hence the solubility of carbon dioxide is greater. On opening the bottle, the pressure reduces to the atmospheric pressure and carbon dioxide gas escapes due to decrease in solubility. The solubility of all gases decreases with rise of temperature.

To compare the solubility of different solids in various solvents, the amount of solid that dissolves in 100 g of the solvent is used. *The solubility of a solid at a given temperature in a given solvent is the number of grams of the solute that will dissolve in 100 g of the solvent in a saturated solution at that temperature.*

9.4 To Determine the Amount of Water of Crystallisation

Experiment 3: Weigh a small, clean and dry porcelain crucible. Powder some crystals of copper sulphate in a mortar and fill about one-

third of the crucible with it. Weigh the crucible again. The difference in the two gives the mass of powdered copper sulphate. Now heat the crucible to drive off the water of crystallisation. Allow it to cool and weigh again.

When copper sulphate crystals are heated, they turn from blue to white. The colour change is not, however, a sure test that all water of crystallisation has been removed, because a few crystals in the lower layers may still contain water. The crucible is, therefore, heated again for about 10 minutes, cooled and weighed. If repeated heating, cooling and weighing gives the same mass, then we can be sure that all the water is driven off.

The percentage of water content may be calculated as follows.

$$\begin{aligned}\text{Let mass of empty dry crucible} &= m_1 \text{ g} \\ \text{mass of crucible+crystal} &= m_2 \text{ g} \\ \text{mass of crucible+anhydrous} \\ \text{residue} &= m_3 \text{ g}\end{aligned}$$

Hence, mass of copper sulphate

$$\text{crystals} = (m_2 - m_1) \text{ g}$$

$$\begin{aligned}\text{mass of water of crystallisation} \\ &= (m_2 - m_3) \text{ g}\end{aligned}$$

$$\begin{aligned}\text{percentage of water of} \\ \text{crystallisation} &= \frac{(m_2 - m_3)}{(m_2 - m_1)} \times 100\end{aligned}$$

9.5 Efflorescence and Deliquescence

Experiment 4: Place small amounts of the following substances in separate watch glasses: calcium oxide (also called quick lime), black copper oxide, anhydrous calcium chloride, sodium hydroxide, magnesium chloride, table salt, washing soda crystals (hydrated sodium carbonate), and hydrated sodium sulphate crystals (also called Glauber's salt).

Weigh each watch glass and set it aside for a few days in an open place. You will observe that the watch glasses containing crystals of washing soda and hydrated sodium sulphate lose weight and become powdery. They lose their

water of crystallisation and hence their crystalline structure (Diagram 9.2). *Substances that lose all or a portion of their water of crystallisation when exposed to dry air are said to be efflorescent and the property is called efflorescence.* This process is aided by increase in temperature and a dry atmosphere.

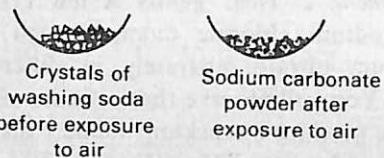


DIAGRAM 9.2 Efflorescence in washing soda.

The watch glasses containing quick lime, anhydrous calcium chloride, black copper oxide, sodium hydroxide, table salt and magnesium chloride show an increase in weight. They form either a solution or they become moist (Diagram 9.3). They absorb water vapour from the atmosphere.

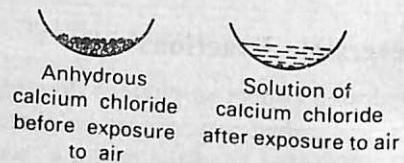


DIAGRAM 9.3 Deliquescence in anhydrous calcium chloride.

Sodium hydroxide and anhydrous calcium chloride absorb water vapour from the atmosphere and form a solution. *Such solids which absorb water vapour from the atmosphere and dissolve in water to form a saturated solution are said to be deliquescent and the property is called deliquescence.* Table salt becomes damp, particularly in humid weather, because of impurities such as magnesium chloride and calcium chloride present in sodium chloride.

Solids like copper oxide and quick lime also absorb water from the atmosphere but they do not form a solution with water. Such solids are said to be *hygroscopic*. The term is also applied to liquids, e.g. concentrated sulphuric acid and

alcohol, which absorb water vapour from the atmosphere.

If a beaker is partly filled with concentrated sulphuric acid and exposed to air for a few days, there is slight increase in the level of the acid due to absorption of water vapour.

Hygroscopy and deliquescence are aided by damp weather and low temperature.

9.6 Drying Agents

Chemists use deliquescent and hygroscopic substances in the laboratory to keep other substances dry. Gases made in the laboratory generally contain water vapour. They are also dried by passing them through either deliquescent or hygroscopic substances.

Solids which need to be kept dry are put in an apparatus called a *desiccator* (Diagram 9.4). The atmosphere inside the desiccator is kept dry by the presence of anhydrous calcium

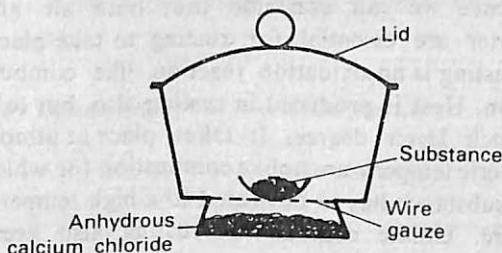


DIAGRAM 9.4 Desiccator.

chloride or concentrated sulphuric acid. Silica gel is another drying agent very often used. A greased lid is used to make the desiccator absolutely air-tight.

Gases are dried in various ways depending on the nature of the gas. The drying agent must be a substance which does not react chemically with the gas to be dried. Usually anhydrous calcium chloride or concentrated sulphuric acid are used for drying gases. For drying ammonia gas an unusual drying agent, calcium oxide (also called quick lime), is used. The apparatus for drying a gas depends upon the choice of the drying agent. Three typical arrangements are shown in Diagram 9.5.

9.7 Rusting

When certain substances are exposed to air they undergo chemical changes. Look at old iron nails, scraps of iron and corrugated iron sheets. You will notice that all these materials, which looked brighter when new, are covered with a reddish brown substance. This substance is called *rust* and this process of corrosion of iron is called rusting.

Rusting is slow oxidation of iron by oxygen of the air in presence of water vapour. Rust is hydrated iron oxide.

Experiment 5: Take some moist iron filings in

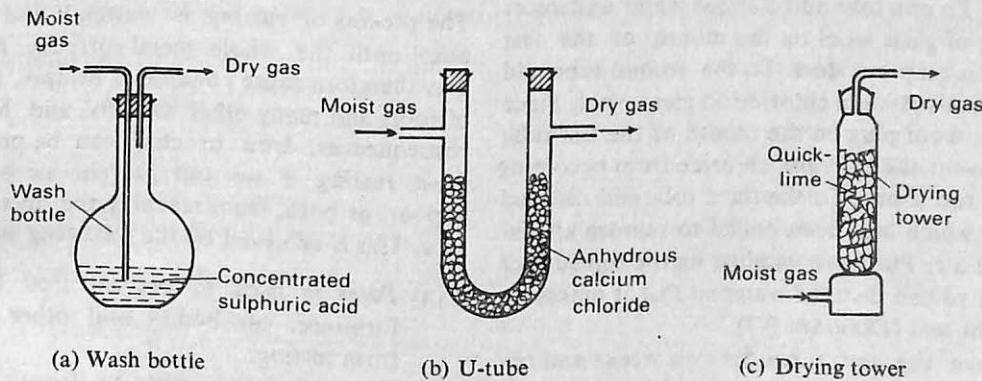
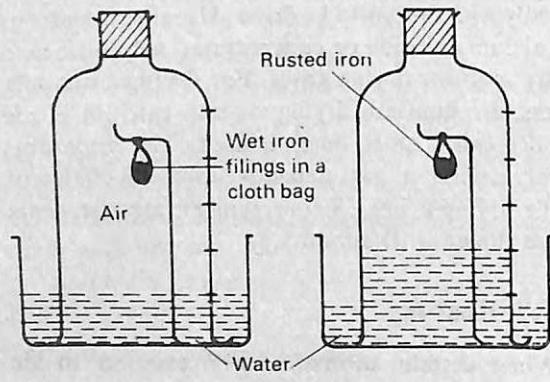


DIAGRAM 9.5 Apparatus used to dry gases.

a bag of cloth. Suspend the bag on a thick iron wire inside a bell jar put in a trough containing water, as shown in Diagram 9.6. The portion of the bell jar above the level of water is divided into five equal parts. Leave the iron filings in the jar for about a week.



(a) Freshly put iron nails (b) After rusting for three or four days

DIAGRAM 9.6 Oxygen is used up during rusting.

After a week you will notice that the iron filings are coated with brown rust and the level of water inside the jar has risen up to the first mark. The gas left is found to be nitrogen. The oxygen has been used up in the process of rusting.

Experiment 6: It can be experimentally shown that both air and water are essential for rusting.

Put new, rust-free iron nails in three test tubes. To one tube add distilled water and insert a plug of glass wool on the mouth of the test tube to keep out dust. To the second tube add anhydrous calcium chloride on glass wool. Place a glass wool plug on the mouth of the test tube to prevent the calcium chloride from becoming moist too soon. To the third tube add distilled water which has been boiled to remove all dissolved air. Put some vaseline on the top surface of the cooled distilled water so that it makes an airtight seal (Diagram 9.7).

Leave the test tubes for two weeks and observe them from time to time. The first tube

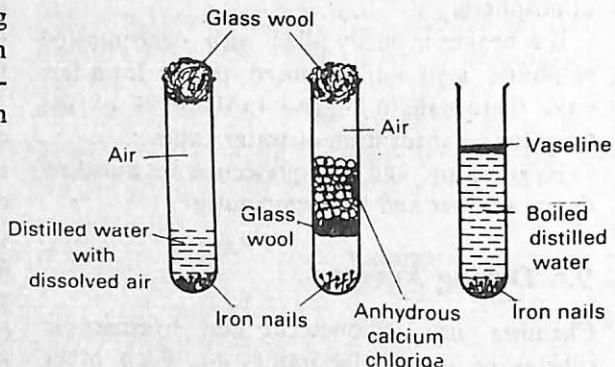


DIAGRAM 9.7 Air and water are essential for rusting.

contains both water and air; you will notice that brown rust is formed on the nails in this test tube. The other two test tubes contain only air and only water, respectively. No rust will be seen to form on the nails in both the tubes. Hence we can conclude that both air and water are essential for rusting to take place. Rusting is an oxidation reaction like combustion. Heat is produced in rusting also, but to a much lesser degree. It takes place at atmospheric temperature unlike combustion for which a substance has to be heated to a high temperature. Unlike combustion, rusting also needs water vapour.

9.8 Prevention of Rusting

The process of rusting is wasteful and continues until the whole metal corrodes. Rusting may therefore cause collapse of bridges, leaking of roofs and many other wasteful and harmful consequences. Iron or steel can be protected from rusting if we can prevent air or water vapour, or both, from reaching the metal surface. This is achieved by the following methods.

- Paint* is used to prevent iron bridges, furniture, car bodies and other objects from rusting.
- Electroplating* is done to deposit a thin

- film of chromium or nickel on the steel surface by *electrolysis*. This film forms a protective coating and prevents rusting.
- (c) Greasing of tools and utensils, which are not regularly used, prevents them from rusting; these objects are rubbed with vaseline or oil.
- (d) *Galvanising* is a process in which corrugated iron sheets and pipes carrying water are coated with zinc by dipping them into baths of molten zinc. The zinc coating prevents iron from rusting.

EXERCISES

1. Solid A is a blue crystalline compound. When heated in a test tube it gives a colourless liquid B and a solid residue C is left in the test tube.
 - (a) Name the substance A.
 - (b) What is the colourless liquid B?
 - (c) Give two tests for liquid B.
.....
.....
 - (d) What is the colour of substance C left in the test tube? Name the substance C.
.....
2. (a) Define water of crystallisation....
.....
.....

(b) Name three compounds having water of crystallisation.
.....

(c) Do all crystals contain water of crystallisation? Explain giving three examples of suitable compounds....
.....
.....
.....
3. Define the following terms.
 - (a) Reversible reaction
.....

(b) Saturated solution.....

(c) Solubility.....
.....
.....

4. What would you observe when the following are exposed to air? Mention also the changes in mass and give reasons for the change.

(a) Quick lime

.....
(b) Concentrated sulphuric acid.....

.....
(c) Sodium hydroxide.....

.....
(d) Washing soda crystals.....

.....
(e) Table salt.....

5. Define the following terms.

(a) Efflorescence.....

.....
.....

(b) Deliquescence.....

.....
.....

(c) Hygroscopy

6. (a) Name three drying agents

.....

(b) Draw the diagram of an apparatus used for drying ammonia gas. Mention the drying agent used.

(c) Can sulphuric acid be used for drying ammonia gas? Give reasons for your

answer.....

.....

.....

.....

7. (a) Give three differences between rusting and combustion (burning).....

.....

(b) Describe an experiment to show that both water and air are necessary for rusting.....

.....

.....

.....

.....

8. How is rusting checked in the following articles?

(a) Iron sheets used for roofing.....

.....

(b) Bicycle parts.....

.....

- (c) Tools.....
- (d) Iron bridges.....

9. Complete the following statements by filling in the blanks.

- (a) During rusting of air is used up.
- (b) Rusting is an reaction because heat is during rusting.
- (c) Water of crystallisation is combination of water molecules with a compound.
- (d) Pure sodium chloride is not a substance because it absorb water vapour from atmosphere when pure.
- (e) Rusting is slow of iron by of the in presence of

Atoms and Molecules

You are familiar with the fact that matter exists in three different forms called the *states of matter*—solid, liquid and gas. Man has always been interested in knowing about the properties of matter. The Indian philosopher *Kanad* gave the name *anu* to the tiny particles which compose all matter. The Greek philosopher *Democritus* suggested that matter is composed of very tiny particles held together by some kind of force. He called these particles *atoms*.

10.1 Dalton's Atomic Theory

John Dalton, in 1803, was the first to use the word atom in the modern sense. His theory may be summarised as follows.

1. The smallest particle of an element is called an atom.
2. An atom can neither be created nor destroyed.
3. All atoms of the same element are alike, but are different from atoms of other elements.
4. Compounds are formed by the chemical combination of atoms in small whole numbers.

Dalton's theory was immediately recognised because it explained the following two laws of chemical combination.

(a) *Law of conservation of mass:* This law states that matter can neither be created nor destroyed in the course of a chemical reaction.

This means that the total mass of the original substances present before a chemical reaction is the same as the total mass of the products formed. According to Dalton's theory, chemical reactions take place by the combination of atoms which can neither be created nor destroyed. Hence there will be no change in mass in the course of a chemical reaction.

(b) *Law of definite proportions or constant composition:* The law states that all pure samples of the same chemical compound always contain the same element chemically combined in a fixed proportion by mass. According to Dalton's theory, compounds are made up of atoms which can neither be created nor destroyed and atoms of an element are exactly alike. Hence different samples of the same compound which has a fixed composition will always contain similar atoms.

chemically combined with fixed masses. The proportion of the masses of elements will also be fixed. For example, black copper oxide can be prepared by different methods. All atoms of copper have the same mass and all atoms of oxygen are similar. Hence the ratio in the masses of copper and oxygen will be same in the different samples. It has been experimentally found that the proportion Cu : O is 4 : 1 by mass in the samples of the oxide.

Discoveries made in the 20th century have led to the modification of the theory, but Dalton's theory is still useful in explaining chemical reactions. Atoms continue to remain the smallest particles responsible for chemical changes.

10.2 Atoms and Molecules

A molecule is the smallest particle of an element or compound which can normally exist separately.

An atom is the smallest particle of an element which takes part in a chemical change and which cannot be further spilt into smaller particles. It may or may not exist separately.

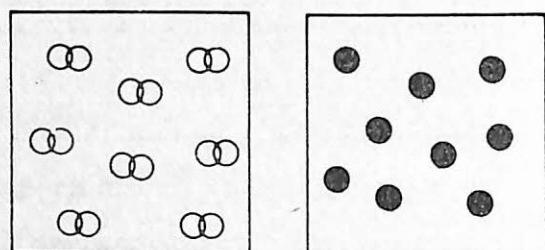
It follows from the definition of an atom that single atoms are not always the smallest particles existing in the sample of an element. Many elements consist of a number of atoms combined together and moving as a single particle. These particles are called molecules.

Most of the elements which are found normally in the gaseous state consist of molecules, each containing two atoms. These are called *diatomic molecules* (Diagram 10.1a). Hydrogen, oxygen, nitrogen, chlorine, bromine (a liquid), and iodine (a solid) form diatomic molecules.

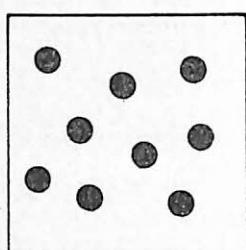
Atoms of helium, neon, argon, krypton and xenon, which are noble gases, exist singly. Their molecules and atoms are similar. Their molecules are thus *monoatomic* (Diagram 10.1b).

A molecule of ozone is *triatomic*; a molecule of phosphorus is *tetratomic*. A molecule of sulphur contains eight atoms. Such molecules are called *polyatomic* molecules.

Atomicity of an element is thus defined as the



○ 1 atom of oxygen
○○ 1 molecule of oxygen



● 1 atom or molecule of helium

(a) (b)

DIAGRAM 10.1 (a) Molecules of oxygen are diatomic

(b) Molecules of helium are monoatomic.

number of atoms present in one molecule of the element.

Since a compound must contain at least two elements, its smallest particle must contain at least two atoms. Hence the smallest particle of a compound can only be a molecule.

10.3 Chemical Symbols

When scientists reached the conclusion that elements are made up of atoms, they tried to find out some means of denoting atoms. Dalton suggested figurative symbols for the atoms of some of the elements as given in Diagram 10.2.



DIAGRAM 10.2 Symbols for atoms suggested by Dalton.

On the basis of these symbol the molecules of water, sulphuric acid and ammonia could be shown as in Diagram 10.3.

This is a tedious method of representing the molecules of compounds and the method was soon abandoned. Berzelius suggested that the elements could be represented by the first letter of the name of the element. Names of some

Water	Sulphuric acid	Ammonia
		
H ₂ O	H ₂ SO ₄	NH ₃

DIAGRAM 10.3 Symbols for molecules of water, sulphuric acid and ammonia according to Dalton's method.

of the elements have the same initial letter. This difficulty was overcome by using the same initial letter, as capital letter, together with one small letter from its name.

Thus a symbol is a short form for the name of an element. It is always represented by a capital letter. *If there are two letters for the symbol of an element, the first letter is a capital letter and the second letter is a small letter.* The following rules are applicable for assigning symbols to different elements.

1. The symbol of an element is the initial letter (capital letter to be used) of the name of the element, e.g.

Carbon has the symbol C

Nitrogen has the symbol N

Hydrogen has the symbol H

2. Some elements have the same first letter; hence a small letter is added to the first letter to avoid any confusion. For example, carbon, calcium and cobalt all have the same initial letter. Their symbols are as follows.

C represents an atom of carbon

Ca represents an atom of calcium

Co represents an atom of cobalt

3. In some cases the symbols of some of the elements are derived from their Latin names. A few examples are given below.

Potassium	Kalium	K
Silver	Argentum	Ag
Mercury	Hydrargyrum	Hg
Gold	Aurum	Au
Tin	Stannum	Sn

The symbol of an element indicates one atom of the element and also its atomic weight. For example, C stands for 1 atom of carbon and means 12 parts by weight on the atomic mass scale.

10.4 Relative Atomic Mass

An atom is a very small particle and cannot be seen even by the most powerful microscope. Physicists have been able to find out the weights of atoms of different elements by indirect methods. An atom of hydrogen weighs 1.67×10^{-24} g and an atom of oxygen weighs 26.5×10^{-24} g. Thus, if we were to write the actual weight of atoms of elements in grams, it would mean working with very small numerical factors.

Just as we have units for length and mass, e.g. a metre and a kilogram respectively, for comparing the lengths and masses of other substances, we could take the mass of some light particle as a unit to compare the mass of atoms.

The hydrogen atom is the lightest atom and was, therefore, chosen as a particle for comparing the masses of atoms of other elements. Hence the atomic mass of an element is more correctly called *relative atomic mass*. *It is defined as the number of times an atom of the element is heavier than the mass of an atom of hydrogen.*

On this standard, the atomic mass of oxygen is almost 16, which means that the mass of an oxygen atom is 16 times the mass of an atom of hydrogen.

In 1961 a different international standard for the unit of atomic mass was adopted. According to this, *1/12 of the mass of an atom of carbon is the unit for comparing the mass of atoms of other elements.* This unit is called atomic mass unit (amu). The atomic masses according to this unit

Element	Latin name	Symbol
Copper	Cuprum	Cu
Lead	Plumbum	Pb
Iron	Ferrum	Fe
Sodium	Natrium	Na

are very similar to the atomic masses when hydrogen is taken as the unit.

10.5 Relative Molecular Mass

The mass of a molecule of an element or a compound is also compared with the mass of a hydrogen atom. It can be defined as the number of times one molecule of the substance is heavier than one atom of hydrogen.

10.6 Isotopes

Atomic masses of different elements have been found by physicists by using an instrument called the *mass spectrograph*. This has led to the discovery of different types of atoms of the same element. These different atoms are called isotopes. *Isotopes are atoms of the same element which have identical chemical properties but have different atomic masses.* There are two types of chlorine atoms, one with atomic mass 35 and the other with atomic mass 37. Normally, they are mixed together in such a proportion that the average atomic mass comes to 35.5.

Table 10.1 gives the symbols and atomic masses of some common elements.

TABLE 10.1: Symbols and atomic masses of some elements.

Element	Symbol	Atomic mass
Aluminium	Al	27
Argon	Ar	40
Barium	Ba	137
Bromine	Br	80
Calcium	Ca	40
Carbon	C	12
Chlorine	Cl	35.5
Chromium	Cr	52
Copper	Cu	64
Fluorine	F	19
Gold	Au	197
Helium	He	4
Hydrogen	H	1
Iodine	I	127
Iron	Fe	56
Lead	Pb	207
Lithium	Li	7
Magnesium	Mg	24
Manganese	Mn	55
Mercury	Hg	201
Neon	Ne	20
Nitrogen	N	14
Oxygen	O	16
Phosphorus	P	31
Potassium	K	39
Radium	Ra	226
Silicon	Si	28
Silver	Ag	108
Sodium	Na	23
Sulphur	S	32
Tin	Sn	119
Uranium	U	238
Zinc	Zn	65

EXERCISES

1. Fill in the blanks in the sentences given below.

- The particle taking part in a chemical reaction is the.....
- A molecule of helium contains.....atom(s) and a molecule of chlorine contains.....atom(s).

(c) Isotopes are atoms of.....element which have.....atomic masses.

(d) Chlorine gas consists of two types of atoms, one with atomic mass.....and the other with atomic mass.....

2. Define the following terms.

(a) Relative atomic mass

.....

(b) Atomicity

.....

(c) Molecule.....

.....

3. Write the symbols of the following elements.

Iron..... Manganese..... Tin..... Lead..... Copper.....

Magnesium..... Mercury..... Silver..... Sodium..... Potassium.....

4. (a) What do you understand by the statement that atomic weight of oxygen is 16?.....

.....

.....

(b) Explain the fractional atomic weight of chlorine.....

.....

.....

(c) List the differences between an atom and a molecule.....

.....

.....

.....

5. State, giving reasons, which of the following statements are true and which are false. Rewrite the false sentences correctly.

(a) All atoms of an element are alike.....

.....

- (b) The total mass of the reactants is the same as the total mass of the product formed in a reaction.....
-
.....

- (c) Copper oxide formed by different methods will have different proportions of copper and oxygen.....
-
.....

6. (a) State Dalton's atomic theory
-
.....
.....
.....
.....

- (b) Explain how any two of his theories have been modified.....
-
.....
.....
.....

11

The Chemical Communication

According to Dalton's atomic theory, molecules of compounds are formed by the combination of atoms of different elements in small whole numbers. Just as a symbol is the short form of an element, a *formula is the short form of a compound in which the symbols of atoms of all the elements present in a molecule of the compound are represented.*

The formula of a compound can be found out experimentally, either by finding out the masses of elements present in the compound or by electrical experiments.

11.1 Empirical and Molecular Formula

Example 1: When magnesium burns in air, it combines with oxygen of the air to form magnesium oxide. In an experiment it was found that 0.26 g of magnesium combined with 0.176 g of oxygen to form magnesium oxide. The formula of magnesium oxide can be calculated from this as below.

	Magnesium	Oxygen
Mass of element	0.26 g	0.176 g

No. of atoms of the element (mass/atomic weight)	$\frac{0.26}{24}$	$\frac{0.176}{16}$
	= 0.011	= 0.011
Dividing by the smallest	$\frac{0.011}{0.011} = 1$	$\frac{0.011}{0.011} = 1$

Hence one atom of magnesium combines with one atom of oxygen. The simplest formula of magnesium oxide is, therefore, MgO.

Example 2: A compound containing sodium, oxygen and carbon when chemically analysed was found to have the following composition: Sodium = 43.4%; Oxygen = 45.3%; Carbon = 11.3%.

Calculate the simplest formula of the compound.

	Sodium	Carbon	Oxygen
Percentage by mass	43.4	11.3	45.3
No. of atoms	$\frac{43.4}{23} = 1.89$	$\frac{11.3}{12} = 0.94$	$\frac{45.3}{16} = 2.83$

Dividing by the smallest	1.89 0.94 = 2	0.94 0.94 = 1	2.83 0.94 = 3
-----------------------------	---------------------	---------------------	---------------------

Hence the simplest formula of the compound is Na_2CO_3 .

The simplest formula of a compound is also called *empirical formula*. It is defined as the formula that shows the number of atoms of different elements in one molecule of the compound in the simplest numerical ratio.

11.2 Molecular Formula

The empirical formula of a compound does not necessarily show the actual number of atoms of different elements present in one molecule of the compound.

For example, the empirical formula of hydrogen peroxide is HO; this means that one molecule of hydrogen peroxide contains equal number of atoms of hydrogen and oxygen. Actually, a molecule of hydrogen peroxide contains two atoms each of hydrogen and oxygen and hence its true formula is H_2O_2 . Glucose is a carbohydrate. Its empirical formula is CH_2O , but its true formula is $\text{C}_6\text{H}_{12}\text{O}_6$.

Thus the true formula or the molecular formula of a compound is the formula that shows the actual number of atoms of different elements present in one molecule of the compound.

In most cases when a compound is formed by the combination of metallic and non-metallic elements, the molecular formula is the same as the empirical formula.

11.3 Radicals

Radicals are groups of atoms of different elements which behave as single units. They combine like atoms, have their own combining power and can be represented by their own formulae. Unlike atoms, however, radicals do not have separate existence of their own. That is, they exist only in combination with some elements or another radical.

For example, sulphate radical is represented by SO_4 , and it contains one atom of sulphur and four atoms of oxygen. It combines with sodium to form a compound called sodium sulphate, Na_2SO_4 . Ammonium radical behaves like a metal and is represented as NH_4 . It contains one atom of nitrogen and four atoms of hydrogen. It can combine with chlorine to form ammonium chloride, NH_4Cl .

11.4 Valency and Formula

Valency of an element or a radical is a measure of its combining capacity. The valency of an element or a radical is the number of hydrogen atoms which can combine with or displace one atom of the element or radical.

For example one atom of chlorine combines with one atom of hydrogen to form hydrogen chloride (HCl); hence the valency of chlorine is one. One atom of oxygen combines with two atoms of hydrogen to form water (H_2O); hence the valency of oxygen is two. One atom of nitrogen combines with three atoms of hydrogen to form ammonia (NH_3); hence the valency of nitrogen is three.

Some elements neither combine nor displace hydrogen. In such cases, we consider the combining power of the element with another element whose valency with hydrogen is known. For example, zinc does not combine with hydrogen but it combines with oxygen, whose valency is two. Since one atom of zinc combines with one atom of oxygen, both must have the same combining power or valency. Hence zinc has a valency of two.

In writing the formula of a compound, we proceed as follows.

1. Write the symbols of elements side by side and their valencies at the top of the symbols.
2. Remove the common factor, if any. Put the number at the top of one element below the symbol of the other element on its right side. Thus the valency numbers are

interchanged. For example the compound of nitrogen and hydrogen is called ammonia and may be written as below:

3 1

NH; hence the formula is N_1H_3

The figure 1 is omitted. Hence the formula of ammonia is NH_3 .

3. If a radical also forms a part of a compound, and if two or more of a radical are needed, the number is indicated by enclosing the radical in brackets and writing the number on their right, outside the brackets; e.g. the valency of zinc (Zn) is two, and the valency of nitrate (NO_3) is one; the formula of zinc nitrate is worked out as below.

2 1

ZnNO_3 ; hence the formula is $\text{Zn}(\text{NO}_3)_2$

4. The metallic part of the compound is always written first, followed by the non-metallic part.

Formula of some compounds are given below.

(a) *Sodium oxide*

1 2

NaO ; hence Na_2O

(b) *Zinc oxide*

2 2

ZnO ; hence ZnO

(c) *Ammonium carbonate*

1 2

NH_4CO_3 ; hence $(\text{NH}_4)_2\text{CO}_3$

(d) *Aluminium sulphate*

3 2

AlSO_4 ; hence $\text{Al}_2(\text{SO}_4)_3$

After getting some practice, you should be able to work this out mentally, and write the formula of the compound in one step. The valency of different elements and radicals are given in the valency chart (Table 11.1).

It is important to remember that the combination of atoms is not always determined by the valencies alone. The formula deduced on the

TABLE 11.1: Valency Chart

Vale- ncy	Metallic	Non-metallic
1	Sodium, Na Potassium, K Silver, Ag Copper I (cuprous), Cu Ammonium, NH_4	Chlorine, Cl Bromine, Br Iodine, I Hydrogen, H Hydroxide, OH Bicarbonate or hyd- rogen carbonate, HCO_3
		Bisulphate or hydro- gen sulphate, HSO_4
		Bisulphite or hydro- gen sulphite, HSO_3
		Bisulphide or hydro- gen sulphide, HS
		Nitrate, NO_3
		Nitrite, NO_2
		Chlorate, ClO_3
2	Magnesium, Mg Calcium, Ca Zinc, Zn Barium, Ba Copper II (cupric), Cu Iron II (ferrous), Fe Lead II (plumbous), Pb Mercury (mercuric), Hg Tin II (stannous), Sn	Carbonate, CO_3 Sulphate, SO_4 Sulphite, SO_3 Sulphur, S Oxygen, O
3	Aluminium, Al Iron III (ferric), Fe	Nitrogen, N Phosphate, PO_4
4	Lead IV (plumbic) Pb Tin IV (stannic), Sn	Carbon, C Silicon, Si

basis of valencies may not exist at all. For example, both sulphur and oxygen are divalent and the formula of the compound formed by the two elements should be SO. In fact no such compound with this composition exists. There are two oxides of sulphur which have the formulae SO_2 and SO_3 , called sulphur dioxide and sulphur trioxide, respectively. Carbon monoxide, CO, should not be possible according to the valency considerations of the two elements; but it exists. Such cases are, however, very few.

In case of combinations between metallic and non-metallic elements or radicals, the valency is a very useful tool for deducing the formula of a compound.

11.5 Number of Atoms in the Formula of Compounds

1. The formula of a compound represents one molecule of the compound. For example the formula of sulphuric acid, H_2SO_4 , denotes one molecule of sulphuric acid containing two atoms of hydrogen, one atom of sulphur and four atoms of oxygen.

2. The formula of copper nitrate, $\text{Cu}(\text{NO}_3)_2$, denotes one molecule of copper nitrate containing one atom of copper, two atoms of nitrogen and six atoms of oxygen. This is because the number 2 outside the brackets means that everything inside the brackets is multiplied by 2.

3. The formula NH_4NO_3 denotes one molecule of ammonium nitrate containing two atoms of nitrogen, four atoms of hydrogen and three atoms of oxygen.

4. If any number is put in front of the formula of a compound, it denotes the number of molecules present and this number multiplies all the figures present in one molecule of the compound. For example, $2\text{H}_2\text{SO}_4$ means two molecules of sulphuric acid containing 4 atoms of hydrogen, 2 atoms of sulphur and 8 atoms of oxygen.

$3\text{NH}_4\text{NO}_3$ means three molecules of ammonium nitrate, containing 6 atoms of nitrogen, 12 atoms of hydrogen and 9 atoms of oxygen.

You should practice the counting of atoms of different elements taking molecules of other compounds.

11.6 Variable Valency

There are certain metals which exhibit more than one valency. When the metal forms a compound exhibiting its lower valency, then 'ous' compounds are formed. When the metal exhibits its higher valency it forms 'ic' compounds.

For example, when a compound of copper is formed in which the metal exhibits valency one (its lower valency), a cuprous compound is formed. When the valency of copper is two (its higher valency), a cupric compound is formed. Similarly in ferrous compounds the valency of iron is two and in ferric compounds the valency of iron is three.

In 1960 the International Union of Pure and Applied Chemistry (IUPAC) abolished the use of 'ous' and 'ic' compounds of metals. It adopted a new system by which the operative valency of the metal forms a part of the name of the compound and is stated in brackets. Some examples are given below.

<i>Old name</i>	<i>New name</i>
Cuprous oxide	Copper (I) oxide, Cu_2O
Cupric oxide	Copper (II) oxide, CuO
Ferrous sulphate	Iron (II) sulphate, FeSO_4
Ferric sulphate	Iron (III) sulphate, $\text{Fe}_2(\text{SO}_4)_3$
Manganese dioxide	Manganese (IV) oxide, MnO_2

You will know the reasons why a metal exhibits variable valency when you study the structure of atoms in higher classes.

11.7 Chemical Equations

A chemical equation shows the results of a chemical change in which the reactants and the products are represented by symbols or formulae.

The reactants are written on the left hand side and the products are placed on the right hand side of the sign '=' or ' \rightarrow '. Thus a chemical equation may be called a short-hand form for a chemical change.

Writing of chemical equations must satisfy certain requirements.

1. The names of the reactants and the products must be established experimentally and must be known.
2. Symbols and formulae of compounds (taking valencies of metallic and non-metallic parts into consideration) must be known.
3. Six elements: hydrogen, oxygen, nitrogen, chlorine, bromine and iodine occur as diatomic molecules. When alone they are written as H_2 , O_2 , N_2 , Cl_2 , Br_2 and I_2 respectively.
4. There must be the same number of each kind of atom on the right side of a chemical equation as on the left hand side, because matter can neither be created nor destroyed in a chemical change (law of conservation of mass). Suitable numbers, if necessary, are inserted before symbols and formulae so that the number of different kinds of atoms in the reactants and products become the same. In most cases, these numbers are small. This step is called *balancing of the equation*.

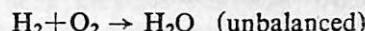
11.8 Examples of Chemical Equations

1. When hydrogen burns in air, it combines with oxygen forming water. The word equation is written as follows:



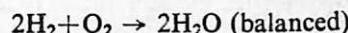
Using symbols for hydrogen and oxygen and writing them as diatomic molecules (because they are in an uncombined form) and writing the formula for water (taking into consideration valencies of hydrogen and oxygen), the skeleton

equation (unbalanced equation) may be written as follows:



We cannot change any of the above symbols or formula.

In the above equation there are two atoms of oxygen on the left hand side and only one oxygen atom in water on the right hand side. If 2 is inserted before hydrogen on the left hand side and 2 is inserted before water on the right hand side, the equation becomes a balanced equation and may be written as below:

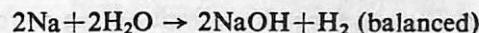


We have 4 atoms of hydrogen and 2 atoms of oxygen on both sides.

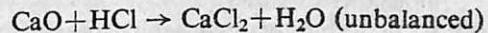
2. Sodium reacts with water to give sodium hydroxide and hydrogen. The skeleton equation is as follows:



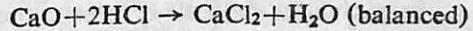
Sodium is balanced but hydrogen and oxygen are unbalanced. If 2 is written before Na, H_2O and NaOH the equation becomes:



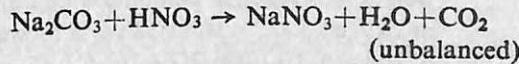
3. When dilute hydrochloric acid reacts with calcium oxide, the products are calcium chloride and water.



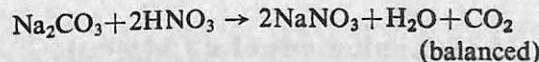
The balanced equation is:



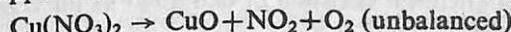
4. When nitric acid reacts with sodium carbonate the products are sodium nitrate, water and carbon dioxide.



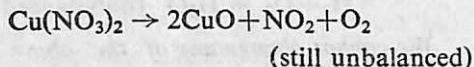
The balanced equation is:



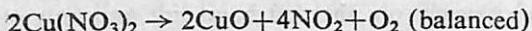
5. When copper (II) nitrate is heated, it gives copper (II) oxide, nitrogen dioxide and oxygen.



For balancing atoms of oxygen the equation may be written as below:



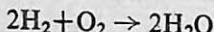
The balanced equation will be:



Balancing an equation is a matter of practice only. Mostly small numbers are involved in balanced equations. Frequent practice will enable you to balance equations without much difficulty.

11.9 Information from a Chemical Equation

A chemical equation provides much information about a chemical reaction. For example, consider the equation:



The equation gives us the following information.

- Hydrogen combines with oxygen to form water.
- Two molecules of hydrogen combine with one molecule of oxygen to form two molecules of water.
- A molecule of hydrogen contains two atoms, a molecule of oxygen contains two atoms, and a molecule of water contains two atoms of hydrogen and one atom of oxygen in combination.
- 4 atoms of hydrogen, i.e. 4 parts by weight of hydrogen combine with 2 atoms of oxygen, i.e. 2×16 or 32 parts by weight of oxygen, to form $2(2+16)$ or 36 parts by weight of water.

11.10 Limitations of a Chemical Equation

A chemical equation does not tell us the following.

- The physical state of the substances, i.e. whether the substances involved are solids, liquids or gases.
- The conditions necessary to bring about a particular chemical reaction. A reaction may be dangerous and precautions may have to be taken for the reaction to take place safely.
- Whether heat energy is evolved or absorbed in the chemical reaction.
- What percentage of the reactants will form the products.
- Whether the reaction is reversible or irreversible.
- How much time is taken for completion of the reaction.

Some of this information is now being included in chemical equations. For example, solid, liquid and gaseous states are indicated by putting in brackets s, l or g (small letters) in front of the symbol or formula of a substance.

Fe(s) is iron in solid state;

Hg(l) is mercury in liquid state, and

CO_2g is carbon dioxide in gaseous state

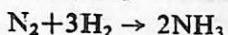
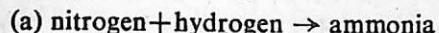
A double arrow (\rightleftharpoons) in between reactants and products represents a reversible reaction. An arrow pointing upwards (\uparrow) indicates a gas being formed, and an arrow pointing downwards (\downarrow) shows the formation of an insoluble substance called a *precipitate*. HCl(aq) means a solution of hydrogen chloride in water; 'aq' stands for aqueous (meaning water).

11.11 Some Typical Chemical Reactions

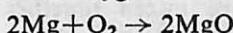
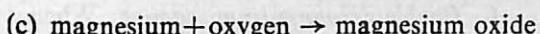
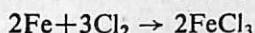
There is a certain pattern of chemical reactions and some principles can be applied to predict the products formed in a chemical reaction. The reactions given below are useful in understanding the nature of chemical reactions.

- Synthesis:* When a compound is formed from its elements or from simpler substances, the

reaction is called synthesis. Some examples are given in the following.

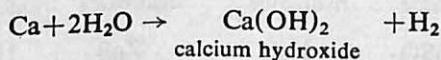
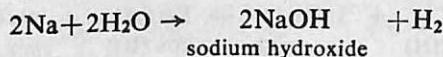


Note the difference between ammonia and ammonium. The former is a compound (a basic gas which turns moist red litmus blue) and the latter is a radical (NH_4) and forms a part of a compound, say ammonium chloride. It cannot exist independently.

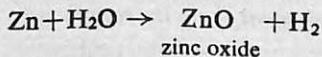
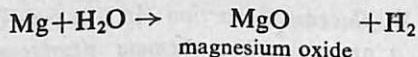


2. Reaction of water with metals: Water reacts with metals under different conditions, depending upon the *reactivity of metals*.

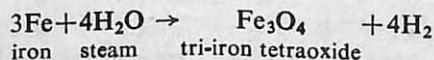
(a) The more reactive metals such as potassium, sodium and calcium react with cold water to produce the hydroxide of the metal and hydrogen.



(b) Hot magnesium and zinc react with hot water or steam to give their oxide and hydrogen.



(c) Red hot iron reacts with steam to give tri-iron tetraoxide.



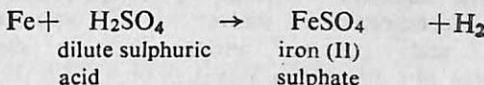
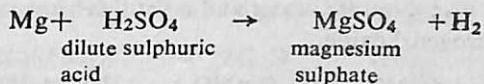
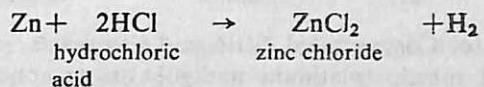
Note that this oxide of iron is neither iron (II) oxide nor iron (III) oxide but is a combination of the two oxides.

Thus we see that the more the metal is reactive the lower is the temperature of water for

the reaction to take place. This confirms that metals have different reactivity. Aluminium, though a very reactive metal, does not react with water because of a compact and rigid deposit of aluminium oxide present on the metallic surface.

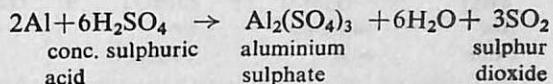
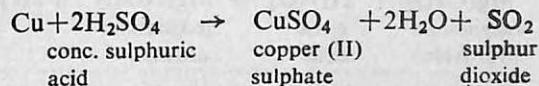
3. Reaction of acids with metals

(a) Dilute hydrochloric acid, concentrated hydrochloric acid and dilute sulphuric acid react with reactive metals (above hydrogen in the reactivity series) to produce salt of the acid and hydrogen.

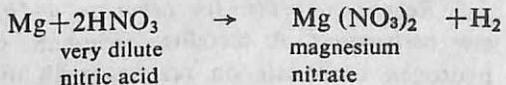


Metals such as copper and silver will not react with the above acids because they are below hydrogen in the reactivity series.

(b) Hot and concentrated sulphuric acid reacts with all metals to produce a sulphate, water and sulphur dioxide. Non reactive metals such as platinum and gold are exceptions.



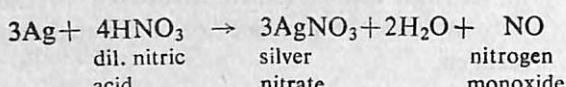
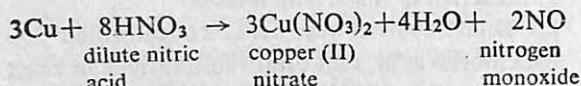
(c) Very dilute nitric acid reacts with magnesium and manganese only to produce a nitrate and hydrogen.



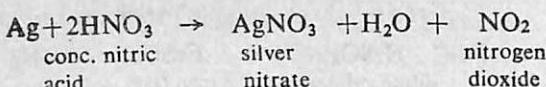
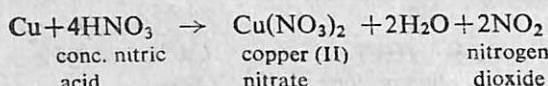


gives the salt of the acid, water and carbon dioxide gas.

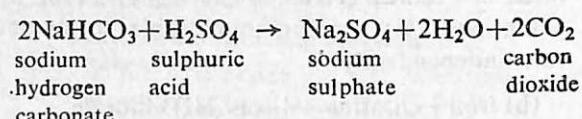
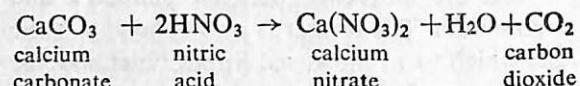
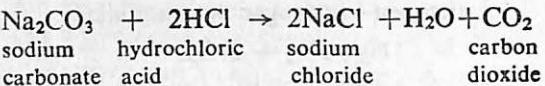
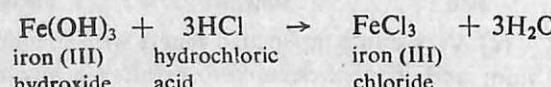
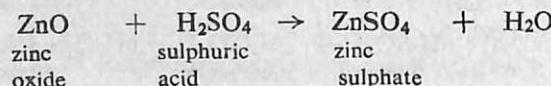
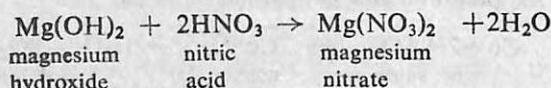
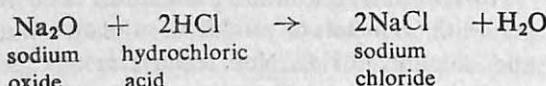
(d) Dilute nitric acid reacts with all metals, except very unreactive metals such as platinum and gold, to give a nitrate, water and nitrogen monoxide (a colourless gas).



(e) Concentrated nitric acid also reacts with all metals (platinum and gold are exceptions) to give a nitrate, water and a reddish-brown gas, nitrogen dioxide.

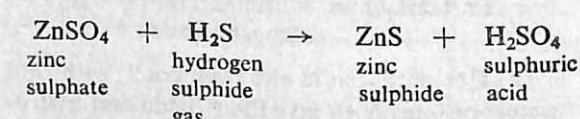
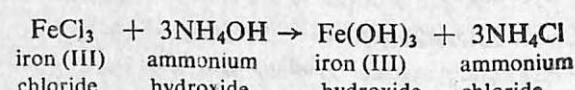
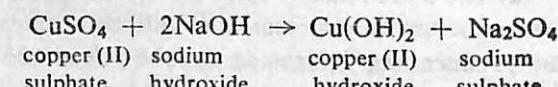


4. Reaction of acids with bases: Metallic oxides and hydroxides are called bases. An acid reacts with a base to give a salt of the acid and water.

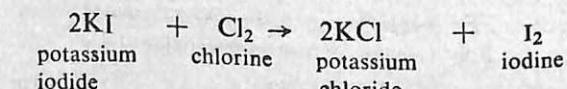
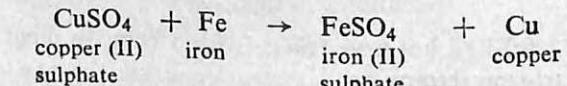


6. Double decomposition reaction: When two compounds are mixed together such that the metallic part of one compound combines with the non-metallic part of the other compound, the reaction is called a double decomposition reaction.

Precipitates from solutions are generally formed by this type of reaction.



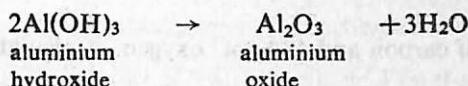
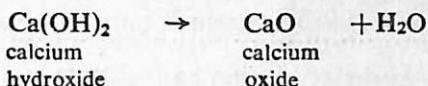
7. Displacement reaction: In this type of reaction, a more reactive element displaces a less reactive element from its compound.



5. Reaction of acids with carbonates and hydrogen carbonates: A metallic carbonate or a hydrogen carbonate on reacting with an acid

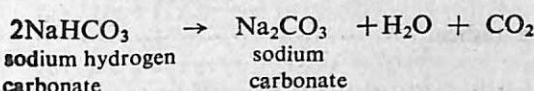
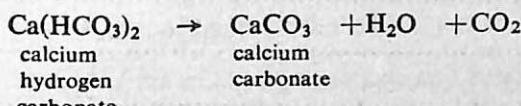
8. Thermal decomposition: In this type of reaction, a compound decomposes into two or more substances on heating.

(a) When metallic hydroxides, except the hydroxides of sodium and potassium, are heated, they produce an oxide and steam.

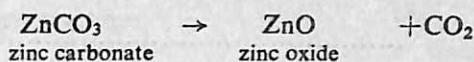
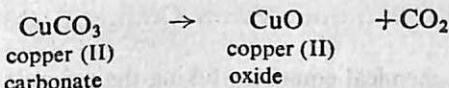


(b) When hydrogen carbonates of metals are heated, the products are a carbonate, water and carbon dioxide.

Hydrogen carbonates of only sodium, potassium, calcium and magnesium exist.

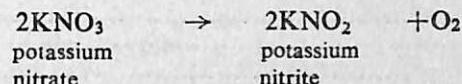


(c) When carbonates of metals are heated, they produce an oxide and carbon dioxide.

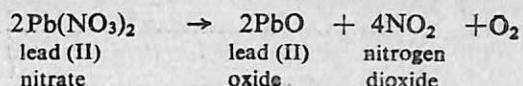
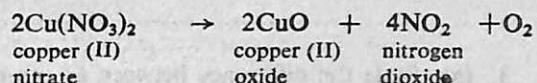


Carbonates of sodium and potassium are stable and do not decompose on heating.

(d) When sodium or potassium nitrate are heated, they decompose at a high temperature to give the nitrite and oxygen.



(e) Nitrates of nearly all other metals decompose on heating to give the oxide, nitrogen dioxide (a reddish brown gas) and oxygen.



Silver nitrate also decomposes in the same manner. But silver oxide further breaks up into silver and oxygen.

EXERCISES

1. Define the following terms.

(a) Empirical formula.....

(b) Molecular formula

(c) Valency.....

2. (a) 5.4 g of an element X combine with 21.3g of chlorine. Find the empirical formula of the metallic chloride. (At. wt. of X = 27, Cl = 35.5)

-
.....
.....
.....
.....
(b) A compound contains 40% of calcium, 12% of carbon and 48% of oxygen. Calculate the empirical formula of the compound. (At. wt. of Ca = 40, C = 12, O = 6.)
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.....
.....

- 3.** **(a)** State the difference between an atom and a radical.....
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- (b)** State what information can be derived from a chemical equation, taking the example of hydrogen combining with chlorine to form hydrogen chloride.....
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.....

- 4.** What information is not conveyed by a chemical equation?
-
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.....

5. Write balanced chemical equations for the following reactions.

- (a) Phosphorus burning in air.....
- (b) Heating of mercury oxide.....
- (c) Heating lead nitrate crystals.....
- (d) Heating copper (II) carbonate.
- (e) Action of hydrochloric acid on iron (II) sulphide.
- (f) Heating of potassium chlorate.
- (g) Decomposition of hydrogen peroxide (H_2O_2) to give oxygen.....
- (h) Sulphur dioxide dissolving in water.
- (i) Action of hydrogen on lead (II) oxide.

6. Complete the following equations and balance them.

- (a) $Na + H_2O \longrightarrow$
- (b) $Fe + Cl_2 \longrightarrow$
- (c) $Cu + H_2SO_4 \longrightarrow$
conc.
- (d) $Fe + H_2SO_4 = \longrightarrow$
dilute
- (e) $Pb + HNO_3 = \longrightarrow$
dilute
- (f) $Zn + HNO_3 = \longrightarrow$
conc.
- (g) $Ca(OH)_2 + HCl \longrightarrow$
- (h) $MgCO_3 + HNO_3 \longrightarrow$
- (i) $CuCl_2 + NaOH \longrightarrow$
- (j) $KBr + Cl_2 \longrightarrow$
- (k) $Fe(OH)_3 \xrightarrow{\text{heated}}$
- (l) $NaHCO_3 \xrightarrow{\text{heated}}$
- (m) $NaNO_3 \xrightarrow{\text{heated}}$

Indicate the type of chemical change that takes place in (b), (g), (i), (j), and (l).

12

Electricity and Chemical Changes

You have studied earlier that electricity passes readily through metallic elements. Copper and aluminium cables are used to bring electricity to our houses. They get heated by the passage of electricity through them, but they do not undergo any chemical change. Such substances are called *conductors*.

Non-metallic elements are bad conductors of electricity. Graphite, a form of carbon is, however, a good conductor of electricity.

12.1 Conduction of Electricity through Compounds

Experiment 1: Connect two platinum wires to a 6-volt battery, with an ammeter reading up to 2 amperes in series (Diagram 12.1). Take distilled water in a beaker. Dip the platinum wires into it and note the current flowing in the ammeter. You will notice that the ammeter does not record any current. This shows that distilled water does not conduct electricity. Repeat the experiment taking the following liquids separately in beakers:

- (a) alcohol
- (b) dilute hydrochloric acid

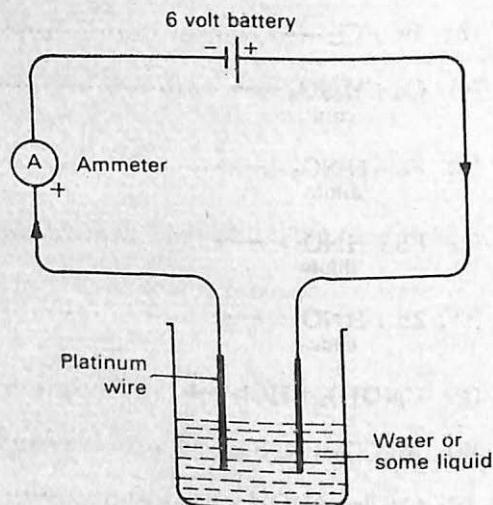


DIAGRAM 12.1 Conduction of electricity by liquids.

- (c) dilute sulphuric acid
- (d) sugar solution in water
- (e) sodium chloride solution in water
- (f) copper sulphate solution in water
- (g) sodium hydroxide solution in water

Remember to wash the platinum wires with distilled water after they have been dipped into

one solution and before dipping into another solution.

You will observe that distilled water, alcohol and sugar solution do not conduct electricity. However, solutions of copper sulphate, sodium hydroxide, sodium chloride, dilute hydrochloric acid and dilute sulphuric acid do conduct electricity. You will also observe that in case of these liquids, some gas is evolved at the platinum wire, showing that a chemical change takes place.

Experiment 2: Connect a steel rod to the negative terminal of a 12-volt battery and a graphite rod to the positive terminal as shown in Diagram 12.2. Fill a crucible, about two-thirds, with solid lead bromide and dip the rods into it. There is no deflection in the ammeter, showing that no electrical current flows through solid lead bromide.

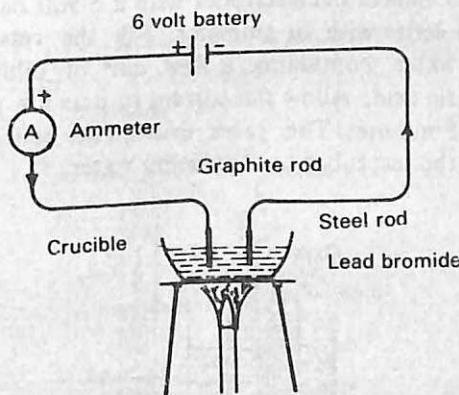


DIAGRAM 12.2 Conduction of electricity through molten lead bromide.

Melt the solid lead bromide by heating it with a bunsen burner. The deflection in the ammeter shows the flow of electric current.

Lead bromide is soluble in water. If rods connected to the battery are dipped into a solution of lead bromide, the flow of electric current is shown by the deflection of the needle in the ammeter.

You will observe that when electricity flows through molten lead bromide, red vapours of bromine are liberated around the carbon rod connected to the positive terminal, and metal lead is deposited around the steel rod connected to the negative terminal of the battery. This clearly indicates the chemical decomposition of the compound which conducts electricity.

12.2 Electrolyte and Electrolysis

As far as the conduction of electricity through compounds is concerned, they can be put into two definite classes. Compounds such as acids, bases and salts conduct electricity and are called *electrolytes*. Compounds such as alcohol and sugar which do not conduct electricity are called *non-electrolytes*. Water conducts electricity at a very high voltage.

Thus *electrolytes* are compounds which conduct electricity, when in molten state or when dissolved in water, and undergo chemical decomposition by the flow of electric current. The process of conduction of electricity through an electrolyte, which undergoes a chemical decomposition by the flow of electricity, is called *electrolysis*.

12.3 Electrodes

The metal or carbon (graphite) rods used in electrolysis are called electrodes. The electrode connected to the *positive terminal* of the battery is called the *anode* while the electrode connected to the *negative terminal* is called the *cathode*.

Platinum and carbon electrodes are generally used in electrolysis because they are non-reactive and hence they react neither with the electrolyte nor with the products of electrolysis.

12.4 Atoms and Ions

Electrolytes contain charged particles, with positive and negative charges, which are attracted by electrodes having opposite charges. Positively charged particles are attracted by the cathode and the negatively charged particles by

the anode. These charged particles are responsible for the conduction of electricity.

Charged atoms or groups of atoms are called ions. In the electrolysis of molten lead bromide, metallic lead is deposited on the cathode (negative plate) and the non-metal bromine is collected around the anode (positive plate). In the compound lead bromide, the lead must, therefore, be a positive ion and bromide must be a negative ion. Experiments on electrolysis of other compounds show that all metallic and hydrogen ions are positively charged and non-metallic ions are negatively charged.

The positively charged metallic and hydrogen ions are called cations. These are discharged at the cathode (negative electrode) during electrolysis. The non-metallic ions are called anions. During electrolysis these are discharged at the anode (positive electrode).

The ions of an element are completely different from their atoms. For example, an atom of sodium is electrically neutral whereas an ion of the element carries a unit positive charge. When sodium metal is added to water, there is a vigorous reaction producing hydrogen gas and sodium hydroxide solution. The sodium ions present in sodium chloride do not react at all with water. Similarly an atom of copper is electrically neutral whereas an ion of copper carries two units of positive charge.

12.5 Charge on Ions

In electrolytes, the ions are responsible for the conduction of electricity. If an electrolyte contains a large number of ions, it is called a *strong electrolyte* and is a better conductor of electricity. On the other hand a *weak electrolyte* contains few ions and is a poor conductor of electricity. Water, ammonium hydroxide and acetic acid are weak electrolytes. Hydrochloric acid, sulphuric acid, nitric acid, the salts of these acids and caustic alkalis are strong electrolytes.

The charge on the ions can be found by applying Faraday's laws of electrolysis, which you will

study in higher classes. *The charge on an ion is the same as the valency of the element or radical.* If an ion is metallic it carries positive charge and if it is non-metallic it carries negative charge. The charge on an ion is indicated on the right of the symbol at the top. Thus CO_3^{2-} indicates that the carbonate ion carries two negative charges. The sodium ion is Na^+ , copper ion is Cu^{2+} , aluminium ion is Al^{3+} and ammonium ion is NH_4^+ . Some of the non-metallic ions are: chloride ion Cl^- , sulphate ion SO_4^{2-} and phosphate ion PO_4^{3-} .

12.6 Electrolysis of Acidified Water

Experiment 3: Different types of apparatus have been used to study the effect of electrolysis of water which is acidified with dilute sulphuric acid. Either platinum or carbon electrodes can be used. In the apparatus shown in Diagram 12.3, two carbon electrodes pass through a vessel. Connect the electrodes with a 6-volt battery in series with an ammeter. Fill the vessel with water containing a few cm^3 of dilute sulphuric acid. Allow the current to pass for at least 15 minutes. The gases evolved are collected in the test tubes by displacing water.

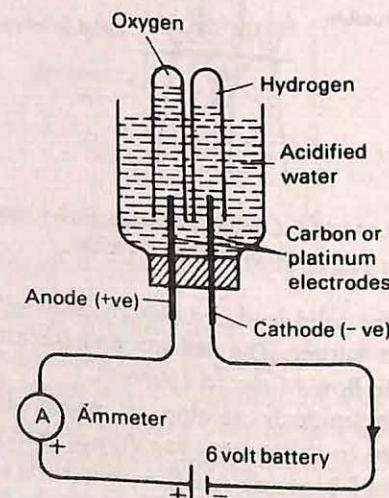
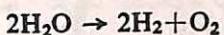


DIAGRAM 12.3 Electrolysis of acidified water.

Compare the volumes of the two gases. Test the gas liberated at the anode with a glowing splint and the gas liberated at the cathode with a burning splint. You will notice that the gas evolved at the anode is oxygen and the gas liberated at the cathode is hydrogen. Oxygen and hydrogen are in the ratio of 1:2 by volume. The products indicate that water has been decomposed into oxygen and hydrogen during electrolysis.



12.7 Electrolysis of Copper Chloride Solution

Experiment 4: Fill a U-tube with 30 per cent solution of copper chloride in water. Dip two platinum electrodes into the solution and connect the electrodes to a 10-volt battery and an ammeter (Diagram 12.4).

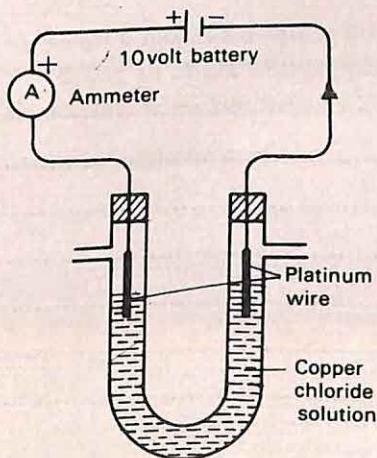


DIAGRAM 12.4 Electrolysis of copper chloride.

You will observe that a reddish-brown coating of copper metal is deposited at the cathode. There is an evolution of a greenish yellow gas around the anode, which comes out of the U-tube and can bleach moist litmus paper. This gas is chlorine.



12.8 Electrolysis of Copper Sulphate Solution

When electricity is passed through copper sulphate solution using both electrodes of copper (Diagram 12.5), the electrolysis does not result in the decomposition of copper sulphate. Copper ions from the copper sulphate solution are deposited on the cathode. At the anode, copper atoms change into copper ions and go into the solution. The reactions at the two electrodes may be written as below:

At cathode : Cu^{2+} changes to Cu
At anode : Cu changes to Cu^{2+}

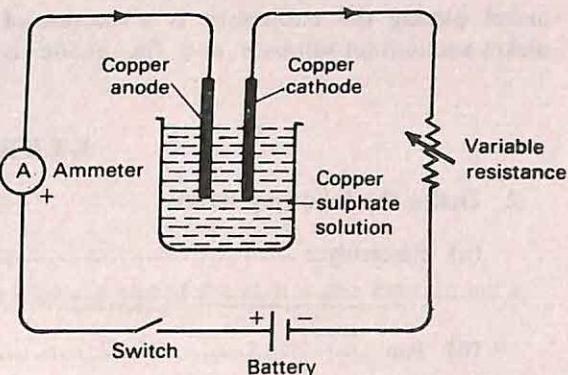


DIAGRAM 12.5 Electrolysis of copper sulphate using copper electrodes.

In effect, copper is transferred from the anode to the cathode. Thus the cathode becomes thicker and the anode becomes thinner as the electrolysis proceeds. This principle is applied in the refining of copper when thin sheets of pure copper are made the cathode and thick sheets of impure copper are made the anode. Pure copper gets transferred from the anode to the cathode and the impurities settle at the bottom of the tank.

12.9 Applications of Electrolysis

1. *Extraction of metals:* More reactive metals such as sodium could be extracted from minerals

only when electrolysis became known to the scientists.

2. *Refining of metals*: A large number of metals are obtained in pure form by electrolysis. Copper refining has been explained earlier.

3. *Electroplating*: It is the process of depositing a thin and compact film of a non-reactive metal on an article made of another metal. Silver and gold plating is done to make an article more attractive. Nickel and chromium plating is done on articles made of steel to prevent rusting. The article to be electroplated is made the cathode (-ve plate). The electrolyte is the solution of a salt of the metal to be deposited, and the anode (+ve plate) is the metal. For nickel plating the electrolyte is a solution of nickel ammonium sulphate and the anode is

made of nickel rod. The object to be electroplated by nickel is made the cathode.

Diagram 12.6 shows a spoon being electroplated with silver.

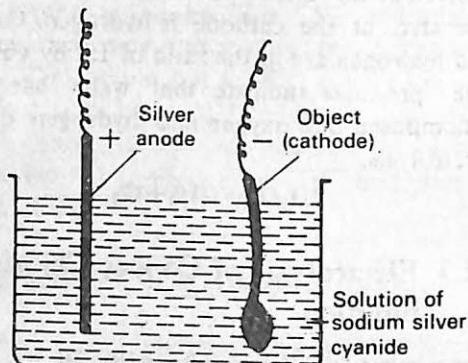


DIAGRAM 12.6 Silver plating.

EXERCISES

2. Define the following terms.

(a) Electrolyte

.....

(b) Ion

.....

(c) Cation.....

.....

(d) Anion.....

.....

2. State the difference between the following.

(a) Conductor and electrolyte

.....

.....

(b) Atom and ion

.....

.....

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.....
.....
- (c) Strong electrolyte and weak electrolyte

3. Fill in the missing words.

Molten lead bromide consists of.....ions and.....ion-s. The ion of.....is called.....and the ion of.....is called..... When electricity is passed through molten lead bromide the.....ions are deposited on the.....and the.....ions are deposited on the..... The electrode connected to the positive terminal is called.....and the electrode connected to the negative terminal is called.....

4. (a) State two industrial uses of

electrolysis

.....
.....

(b) Draw a labelled diagram for electro-plating an object with copper.

5. (a) Explain what you would observe when electricity is passed through water containing a little sulphuric acid

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(b) How is pure copper obtained from impure copper?.....

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6. (a) Name an acid which is a weak electrolyte.....

(b) Give one example each of an ion containing more than one element, and an ion having a positive charge.....

7. Explain the following.

(a) Solid sodium chloride does not conduct electricity whereas molten sodium chloride does.....

(b) An aqueous solution of acetic acid is a poor conductor of electricity whereas dilute sulphuric acid is a good conductor.....

8. (a) Write down the ions present, along with the charges on them, in sodium carbonate, ammonium sulphate and sulphuric acid.....

.....
.....

(b) State what would happen when electricity is passed through molten magnesium chloride.

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Chemistry of Rocks and Minerals

13.1 Minerals and Ores

The earth is a treasure house of chemical materials. These materials generally contain various elements and compounds which are called *minerals*. These minerals are mixed with impurities such as clay and sand. The mixture of impurities and minerals is called *ore*. Metals are extracted from ores. Haematite, an ore containing the mineral iron (III) oxide, is used to obtain iron and steel. Bauxite is an ore of aluminium, and is used in the extraction of the metal. Some unreactive metals, such as gold, are found in the free state. Sulphur also occurs in the free state.

13.2 Sources of Calcium Carbonate

Calcium carbonate is abundantly found in the form of rock formations such as chalk, limestone and marble. The mineral *Iceland Spar*, also called *calcite*, is a very pure form of calcium carbonate. (The chalk mentioned here is very different from the blackboard chalk which is mainly calcium sulphate.)

Some sea organism such as corals and molluscs take in calcium compounds dissolved in

sea water to build their shells and other hard parts. When these animals die, they sink to the bottom of the water and decompose. Their hard parts and shells remain deposited on the sea bed. This deposit becomes thicker over a period of thousands of years. Owing to the upheavels of the earth's crust, these deposits are raised and become *chalk hills*. High pressures due to other deposits and high temperatures inside the earth, change some chalk deposits to *limestone*. Some rocks, however, are subjected to very high temperature and pressure; these change to *marble*. Thus, chalk, limestone and marble are different stages of change taking place in the same chemical substance, calcium carbonate. Chalk is soft and powdery, limestone is hard and stony, while marble is very hard and crystalline.

13.3 Action of Heat on Chalk or Limestone

Experiment 1: Heat strongly some powdered chalk or limestone in a hard glass test tube. Pass the gas liberated into a test tube containing lime water (Diagram 13.1). The lime water

turns milky showing that carbon dioxide gas is given out.

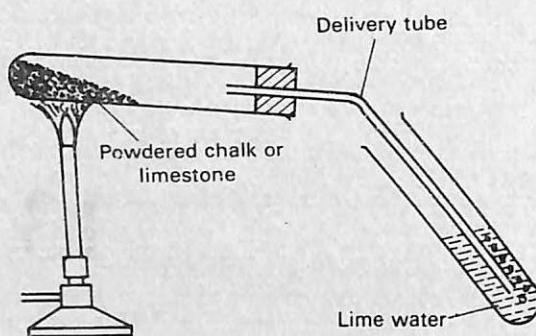
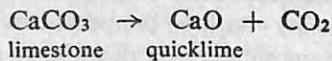
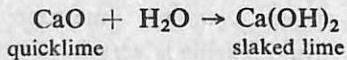


DIAGRAM 13.1 Carbon dioxide is evolved on heating powdered chalk or limestone.

Allow the residue in the test tube to cool. Transfer it into an evaporating dish. Add a few drops of water on to the solid. You will notice that the solid crumbles to powder, swells up and becomes very hot. You may also observe steam coming out. Calcium carbonate in any form, on heating strongly, changes to calcium oxide, called *quicklime*. Carbon dioxide gas is liberated.



The white powder obtained when quicklime is treated with a small amount of water is chemically calcium hydroxide and is called *slaked lime*.



Experiment 2: Take a small amount of slaked lime in a test tube, add about half a test tube full of distilled water and shake well. Filter and test the filtrate with red litmus paper. The filtrate turns red litmus blue showing that calcium hydroxide is slightly soluble in water and the solution is basic. Most of the calcium hydroxide does not dissolve in water. The solution of calcium hydroxide in water is called *lime water*.

13.4 Uses of Limestone

1. Large quantities of quicklime and slaked lime are prepared by strongly heating limestone and chalk in lime kilns (Diagram 13.2). The heating produces quicklime which is then converted to slaked lime by treatment with water.

The soft paste made by mixing slaked lime, sand and water is called *lime mortar*, and is used for holding bricks together in building structures.

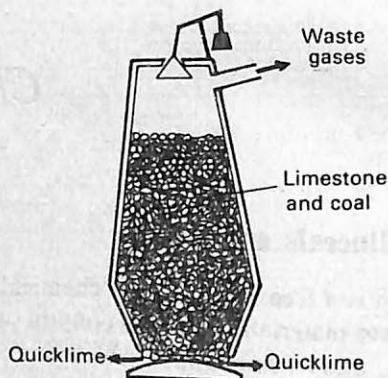


DIAGRAM 13.2 Manufacture of quicklime from limestone or chalk.

2. A thick suspension of slaked lime in water is used for white washing buildings.

Both these uses are based on the fact that calcium hydroxide combines with carbon dioxide of the air to form a hard layer of calcium carbonate. Slaked lime is used in agriculture for removing acidity of the soil and for better drainage through clay soils.

3. *Cement* is made by mixing limestone and clay. A finely powdered mixture of the two is mixed with water to form a sludge or slurry which is heated to about 1500°C in a long rotary kiln. The resulting substance is a mixture of calcium silicate and aluminium silicate (complex substances). Calcium sulphate (about 4%) in the form of gypsum is added to the cooled mass to regulate the setting time. The whole mass is then ground to powder which is grey in

colour. This is called cement or portland cement. When mixed with water, it sets into a hard mass. Cement when mixed with gravel and sand is called *concrete* and when steel rods are embedded into concrete, it is called *reinforced concrete*.

13.5 Preparation of Carbon Dioxide

You have learnt that carbon dioxide is constantly added to the atmosphere due to burning of carbon containing compounds and by respiration in living beings. This gas is at the same time being used up by plants in the process of *photosynthesis*, thus maintaining the amount of carbon dioxide in the atmosphere fairly constant.

Experiment 3: Add dilute hydrochloric acid through a thistle funnel to marble or limestone chips taken in a flask. There is rapid effervescence. Do not collect the gas immediately as it will be mixed with air. Carbon dioxide is heavier (denser) than air. Hence it is collected by upward displacement of air, i.e. by downward delivery (Diagram 13.3). A solution of calcium chloride is left in the flask.

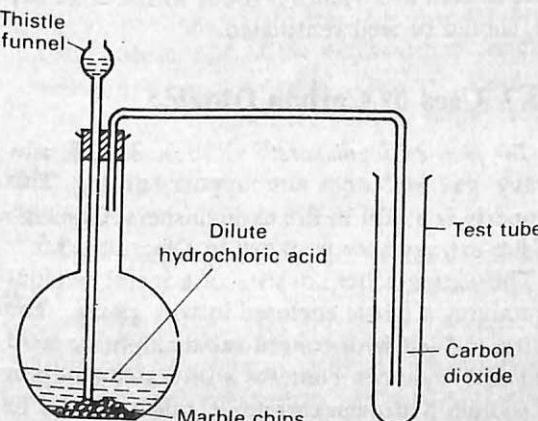
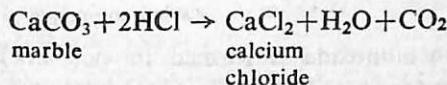
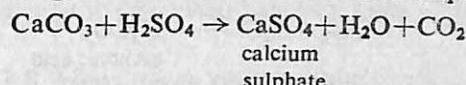


DIAGRAM 13.3 Preparation of carbon dioxide

Dilute sulphuric acid cannot be used in this experiment because the reaction stops after a while owing to the formation of insoluble calcium sulphate which covers the marble chips.



Dilute nitric acid can, however, be used instead. Carbon dioxide is not very soluble in water; it can, therefore, be collected by downward displacement of water if it is not needed dry.

13.6 Properties of Carbon Dioxide

Physical Properties

1. Carbon dioxide is present in air in small amounts. The gas is colourless and odourless.
 2. Introduce a lighted splint into a test tube filled with the gas; you will notice that the splint is extinguished. The gas neither burns nor supports combustion.

Chemical Properties

1. Invert a test tube filled with carbon dioxide into a beaker containing water. After a few minutes you will notice that the water level rises slightly into the test tube. This shows that carbon dioxide is slightly soluble in water. (Diagram 13.4).

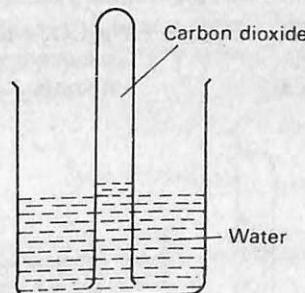
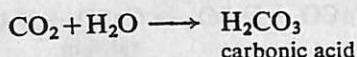


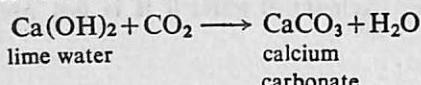
DIAGRAM 13.4 Carbon dioxide is slightly soluble in water.

2. Bubble carbon dioxide for some time into a test tube containing distilled water, and

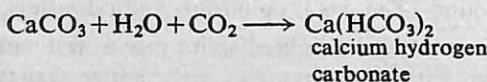
then add a drop of blue litmus solution to it. The litmus turns dull red. The gas when dissolved in water forms a weak acid called carbonic acid.



3. Bubble carbon dioxide gas into lime water for some time. You will notice millessness which is due to the formation of white insoluble particles of calcium carbonate.



The milkiness disappears, when more of carbon dioxide is passed. Due to the formation of calcium hydrogen carbonate, which is soluble in water, a colourless solution is obtained.



4. Invert another test tube filled with carbon dioxide into a beaker containing sodium hydroxide solution. Shake the test tube keeping it immersed in the solution. You will notice that the solution of sodium hydroxide rises appreciably into the test tube (Diagram 13.5). Carbon dioxide reacts with sodium hydroxide to form sodium carbonate. With excess of carbon dioxide, sodium hydrogen carbonate is formed. A concentrated solution of sodium hydroxide is used to absorb carbon dioxide gas.

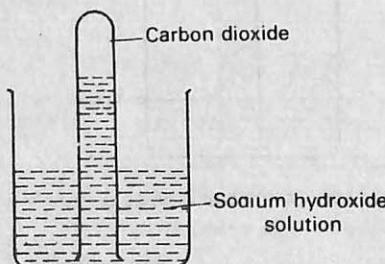
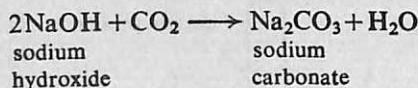
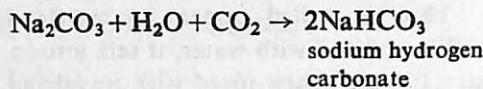
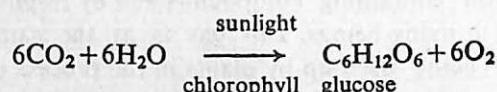


DIAGRAM 13.5 Carbon dioxide reacts with sodium hydroxide.



Both sodium carbonate and sodium hydrogen carbonate are fairly soluble in water.

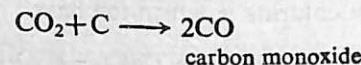
5. Carbon dioxide is taken in by the leaves of plants and water is taken in by the roots. The two combine in the presence of chlorophyll (catalyst) in the leaves, and sunlight (source of energy), to produce carbohydrates, e.g. glucose. This process is known as *photosynthesis*.



The oxygen liberated in photosynthesis keeps the percentage of oxygen in the atmosphere constant.

Photosynthesis differs from respiration. The latter takes place all the time in living cells using oxygen and giving out energy. Carbon dioxide and water are produced in respiration.

6. Carbon dioxide, on passing through a red hot layer of coke, produces carbon monoxide gas, which is extremely poisonous, and burns with a blue flame.



Carbon monoxide is formed in this way in a coke or coal fire. Hence a room with a coke fire on, should be well ventilated.

13.7 Uses of Carbon Dioxide

1. In fire extinguishers: Carbon dioxide is a heavy gas and does not support burning. This property is useful in fire extinguishers. One type of fire extinguisher is shown in Diagram 13.6.

The extinguisher consists of a metal cylinder containing a bottle enclosed in wire gauze. The bottle is filled with concentrated sulphuric acid and the cylinder contains a saturated solution of sodium hydrogen carbonate filled almost to the top. When the knob is struck on the ground, the bottle breaks. The reaction between the

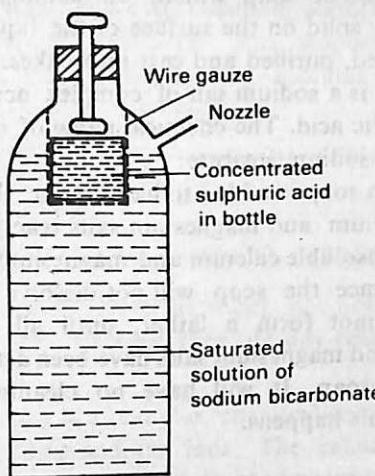
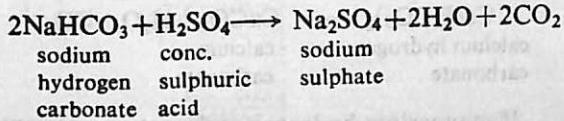


DIAGRAM 13.6 Fire extinguisher.

acid and sodium hydrogen carbonate produces carbon dioxide. The pressure of the gas forces a jet of foamy water through a nozzle. Carbon dioxide, being a dense gas, forms a blanket around the fire and cuts off the supply of air to the fire. The foamy water does the extinguishing by cooling the burning material.



This type of fire extinguisher is not safe for putting out electrical fires, because the liquid which comes out of the extinguisher conducts electricity and may cause serious accidents. Modern fire extinguishers contain liquid carbon dioxide which when directed on to a fire cuts off the supply of air, and also cools the burning material.

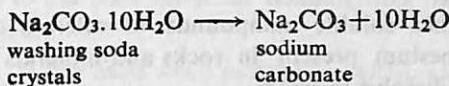
2. *In refrigeration:* Solid carbon dioxide, called *dry ice*, is used as a cooling agent.

3. *In drinks:* Carbon dioxide under high pressure is more soluble in water. Its solution in water is called soda water. Lemonade, Campa-cola, etc. contain carbon dioxide dissolved under a pressure of about 8 atmospheres.

4. *In making cakes and bread:* Cakes and bread swell up forming hollow spaces because of carbon dioxide produced from baking powder, a mixture of sodium hydrogen carbonate and a weak crystalline acid, e.g. tartaric acid.

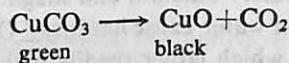
13.8 Heating of some other Carbonates

1. *Washing soda:* This is a white crystalline hydrated compound. When heated in a test tube it loses its water of crystallisation completely and becomes a white anhydrous powder. Drops of water deposit on the cooler parts of the test-tube.

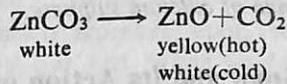


There is no further decomposition when the white powder is heated to a higher temperature.

2. *Copper carbonate:* This is a green powder. When heated in a test tube it gives out carbon dioxide gas which is colourless and odourless and turns lime water milky. A black residue of copper (II) oxide is left in the test tube.



3. *Zinc carbonate:* This is a white powder. When heated in a test tube it gives out carbon dioxide gas, a colourless and odourless gas which turns lime water milky. The residue left is of zinc oxide which is yellow when hot and white when cold.

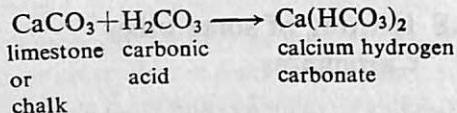


13.9 Action of Water on some Rocks and Minerals

1. Calcium carbonate occurs in nature as limestone and chalk, which are porous to water. During rainfall the carbon dioxide present in air forms a solution of carbonic acid. This reacts

with the limestone or chalk present in rocks forming a solution of calcium hydrogen carbonate. Hence, water collected from places which have limestone or chalk deposits contains dissolved calcium hydrogen carbonate.

The reaction can be written as follows:



2. Rain water seeping through the soil also dissolves small quantities of various minerals. If this water comes in contact with gypsum, some of the mineral will dissolve because calcium sulphate present in it is slightly soluble in water.

Other soluble compounds of calcium and magnesium present in rocks and minerals will also dissolve in water.

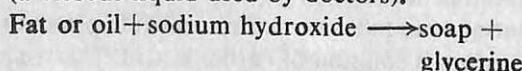
13.10 Hard and Soft Water

You must have experienced that sometimes water requires a lot of soap before a lather is formed. Such water, usually from a spring or a well, which does not lather easily with soap, is said to be hard water. Calcium and magnesium salts dissolved in water cause hardness. Distilled and rain water, which lather readily with soap, are said to be soft water.

There are two types of hardness in water. Temporary hardness which can be removed by boiling, is due mainly to the presence of calcium hydrogen carbonate. Permanent hardness, which cannot be removed by boiling, is due mainly to the presence of calcium sulphate.

13.11 Soap and Its Action on Hard Water

Soap is prepared by heating vegetable oil or animal fat with sodium hydroxide solution. The reaction liberates soap along with glycerine (a viscous liquid used by doctors).



Sodium chloride is added to assist in the separation of soap which, on cooling, forms a hard solid on the surface of the liquid. It is separated, purified and cast into cakes.

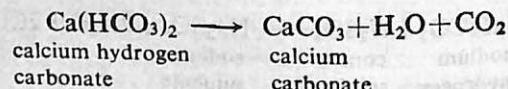
Soap is a sodium salt of complex acids such as stearic acid. The chemical name of common soap is sodium stearate.

When soap is added to hard water the soluble calcium and magnesium salts react to form white insoluble calcium and magnesium stearates. Hence the soap will not dissolve in water and cannot form a lather, until all the calcium and magnesium salts have been acted upon by the soap. It will have no cleaning effect until this happens.

13.12 Methods of Changing Hard Water into Soft Water

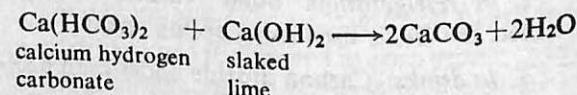
(a) Temporary Hardness

1. When water containing calcium hydrogen carbonate is boiled, insoluble calcium carbonate is formed which can be removed by filtration or decantation. The decomposition takes place as follows:



If magnesium hydrogen carbonate is present, it will change and form a precipitate of magnesium carbonate in a similar manner.

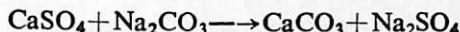
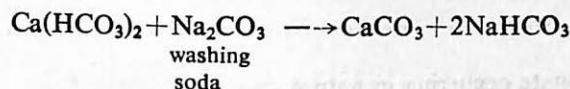
2. By adding just enough amount of slaked lime, all the hydrogen carbonate changes to insoluble carbonate which can be removed by filtration or decantation. It should be remembered that any excess of slaked lime will itself cause hardness.



The second method is cheaper and therefore preferred to the first for removing temporary hardness.

(b) Both Permanent and Temporary Hardness

1. Sodium carbonate, i.e. washing soda, reacts with hydrogen carbonates and sulphates of calcium or magnesium, forming insoluble carbonate which can be removed.



Sodium salts present in water do not cause any hardness because they do not react with soap.

2. Permutit is a complex chemical compound which contains sodium ions. The calcium or magnesium ions present in hard water can get attached to the permutit and the sodium ions present in the permutit get dissolved in water (Diagram 13.7). Thus water becomes free from calcium and magnesium ions which cause hardness.

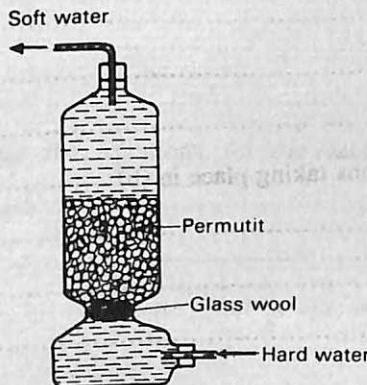


DIAGRAM 13.7 Removal of hardness of water by permutit.

After a time, there are no sodium ions left in the permutit. Hence it becomes unfit for

changing hard water to soft water. The permutit is 'reactivated' by passing a concentrated solution of sodium chloride through the spent permutit.

13.13 Disadvantages of Hard Water

(a) The cleaning action of soap depends upon the formation of lather. A lot of soap is wasted to remove all the calcium and magnesium ions present in hard water as a curdy scum. Modern detergents such as 'Surf' or 'Det' do not react with calcium and magnesium ions in hard water; hence the problem of soap wastage is less in laundry work. But the problem arises if we have to use hard water for washing our hands and face.

(b) When hard water is used in kettles and in boilers, calcium carbonate is deposited as a hard 'fur' or 'scale' inside these. The fur is a bad conductor of heat; hence, the efficiency of boilers or kettles in transferring heat energy is much reduced, leading to wastage. Fur deposits may also lead to bursting of boiler pipes.

13.14 Advantages of Hard Water

(a) Water is carried to our homes in lead pipes. Soft water can dissolve lead and form the slightly soluble lead hydroxide which is a poisonous substance. Hard water does not dissolve lead. Hence hardness in water sent to our homes has an advantage.

(b) Human bones are composed of calcium phosphate and calcium carbonate. Proper growth of bones is dependent upon the calcium compounds in our diet. When we drink water having a certain degree of hardness, the calcium salts present in water help in the formation of strong bones and teeth.

E X E R C I S E S

1. (a) What is the difference between an ore and a mineral?.....

.....

(b) Name the three forms of calcium carbonate occurring in nature

.....

2. (a) Describe all that you would observe when calcium oxide is treated with water.....

.....

(b) How would you prepare quicklime, slaked lime and lime water, starting from limestone?

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(c) Write equations wherever possible, of reactions taking place in (b)

.....

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3. (a) State two uses of lime

.....

.....

(b) How is cement prepared?.....

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4. (a) Draw a labelled diagram for the preparation of carbon dioxide in the laboratory.
- (b) Write the equation for the reaction.

.....

- (c) Can sulphuric acid be used in the laboratory preparation? Why?.....

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5. (a) Describe all that you would observe when carbon dioxide is bubbled through lime water.

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- (b) Write equations for the reactions taking place in (a)

.....

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- (c) Write the equations for the reaction when carbon dioxide is bubbled through caustic potash.....

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- (d) Why do we not use water or lime water for absorbing carbon dioxide?

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6. (a) State the importance of carbon dioxide found in air.

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- (b) State briefly the functioning of a fire extinguisher. Give an equation for the reaction.

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(c) List the uses of carbon dioxide.....

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7. Describe all that you would observe when

(a) Zinc carbonate is heated

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(b) Copper carbonate is heated.....

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(c) Washing soda crystals are heated.....

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.....

8. (a) Why are chalk hills eroded faster than other rocks? What is the consequence of this?

Give an equation for the reaction involved

.....
.....

(b) Mention two compounds which cause temporary hardness and two compounds which cause permanent hardness.....

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(c) Define hardness of water.....

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9. (a) Why is only temporary hardness removed by using slaked lime?

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(b) Describe the method by which both temporary and permanent hardness can be removed using washing soda

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10. (a) What is the principle of functioning of permuntit in changing hard water to soft water?

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(b) Mention two advantages and two disadvantages of hard water.....

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11. (a) How is soap prepared?

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(b) Discuss the action of soap on hard water.....

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- (c) List the uses of carbon dioxide.....
.....
.....
7. Describe all that you would observe when
- (a) Zinc carbonate is heated.....
.....
- (b) Copper carbonate is heated.....
.....
- (c) Washing soda crystals are heated.....
.....
.....
8. (a) Why are chalk hills eroded faster than other rocks? What is the consequence of this?
Give an equation for the reaction involved.....
.....
.....
- (b) Mention two compounds which cause temporary hardness and two compounds which
cause permanent hardness.....
.....
- (c) Define hardness of water.....
.....
9. (a) Why is only temporary hardness removed by using slaked lime?
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- (b) Describe the method by which both temporary and permanent hardness can be removed using washing soda
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10. (a) What is the principle of functioning of permuntit in changing hard water to soft water?

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(b) Mention two advantages and two disadvantages of hard water.....

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11. (a) How is soap prepared?

.....

.....

(b) Discuss the action of soap on hard water.....

.....

.....

14

Acids and Their Salts

In this chapter we shall study how the three common acids and their salts can be prepared in the laboratory.

14.1 Hydrogen Chloride Gas

Experiment 1: Place some sodium chloride in a flat flask and add concentrated sulphuric acid through a thistle funnel. Vigorous effervescence takes place. Hydrogen chloride gas is evolved which is collected by upward displacement of air as the gas is heavier than air. It is dried by bubbling through concentrated sulphuric acid (Diagram 14.1).

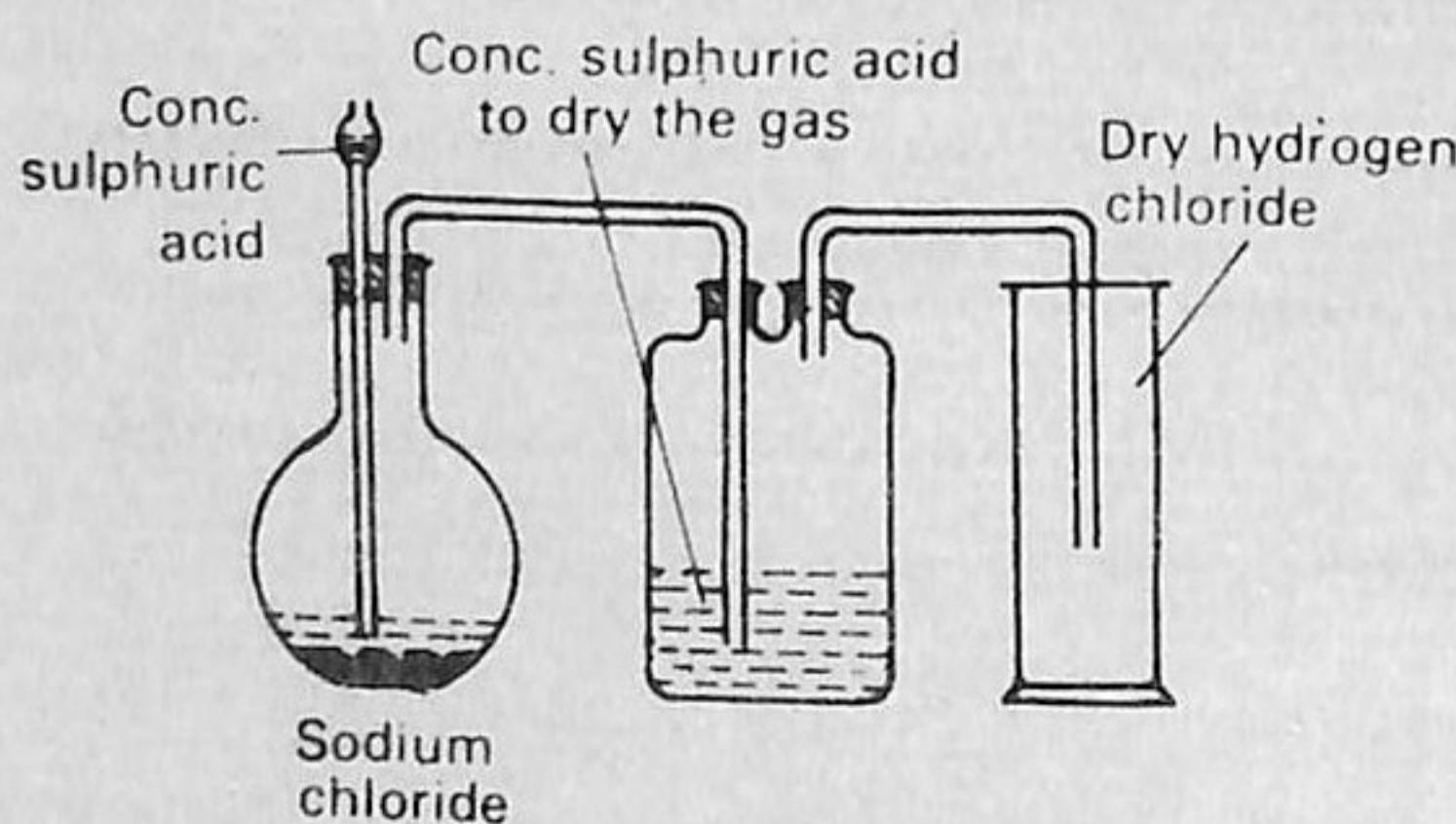
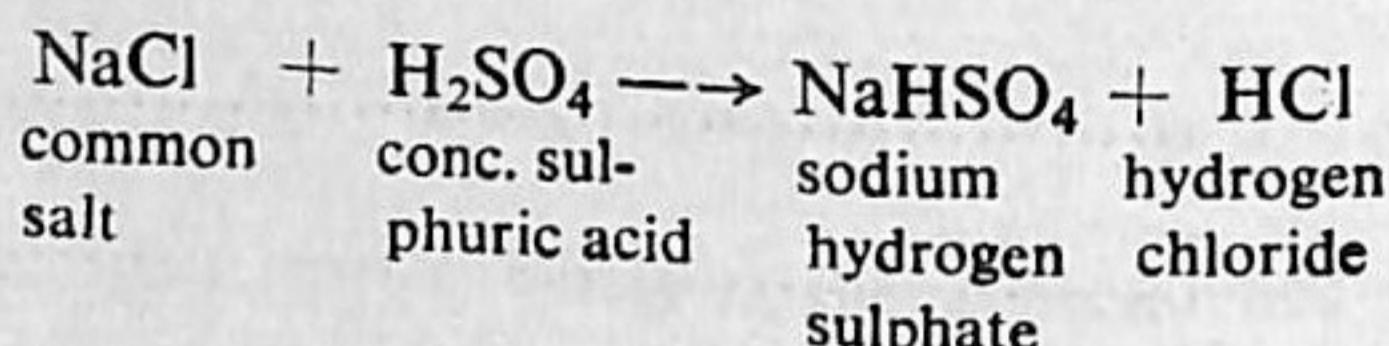
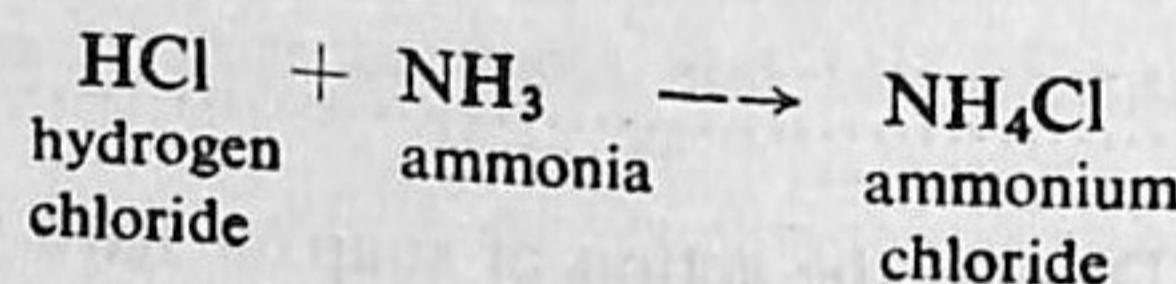


DIAGRAM 14.1 Preparation of hydrogen chloride gas.

A solution of sodium hydrogen sulphate is left in the flask.



Hydrogen chloride gas fumes in air because it forms droplets of hydrochloric acid with the water vapour of the atmosphere. When a rod dipped in ammonia solution is introduced into the gas, dense white fumes of ammonium chloride are produced.



A saturated solution of hydrogen chloride is formed by connecting the delivery tube to an inverted funnel, whose rim just dips in water (Diagram 14.2).

If the delivery tube is immersed in a beaker containing water, the liquid is sucked back into the flask. This is due to the fact that hydrogen chloride dissolves faster in water than it is produced by the reaction.

A concentrated solution of hydrogen chloride in water is concentrated hydrochloric acid.

Hydrogen chloride gas

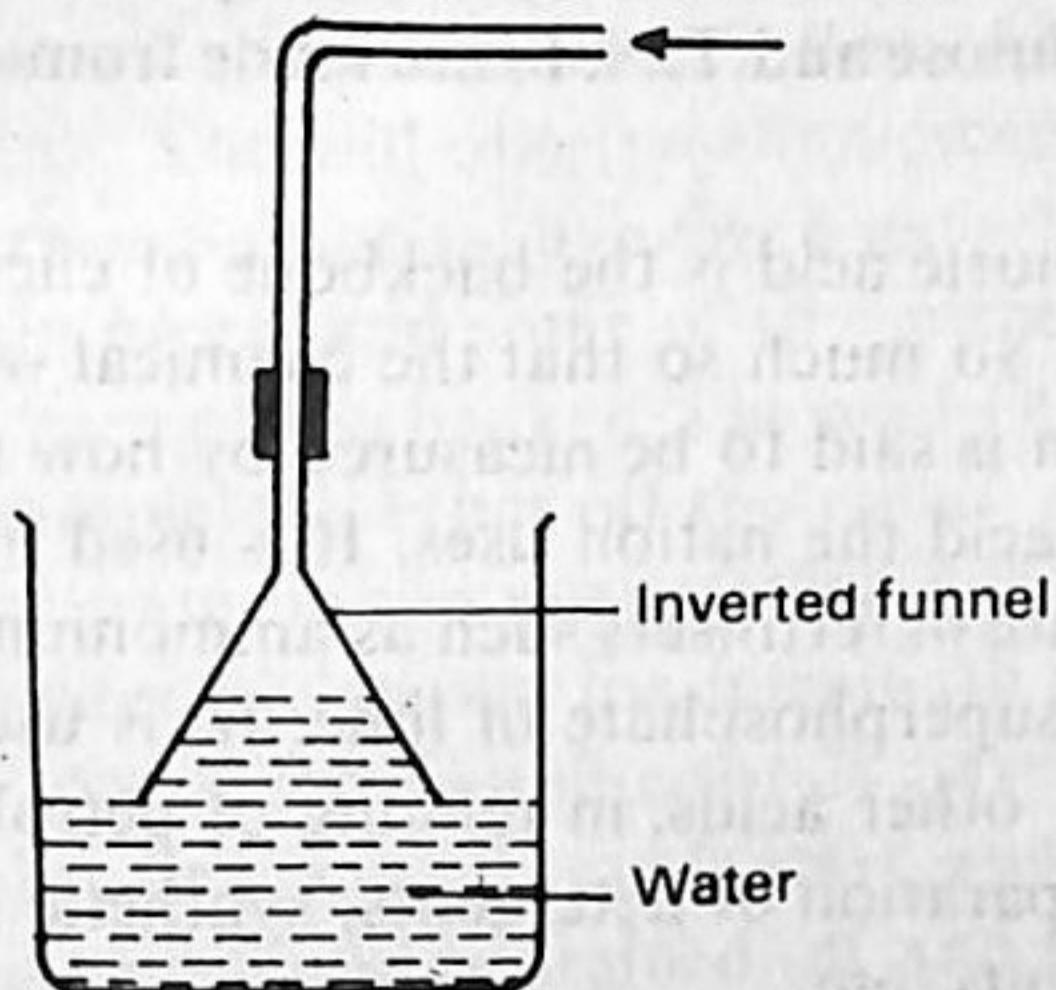
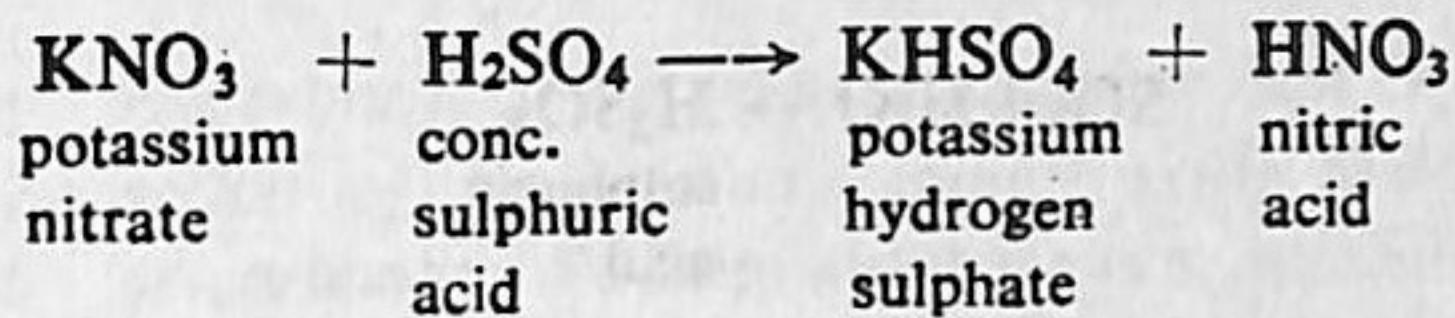


DIAGRAM 14.2 Formation of hydrochloric acid from hydrogen chloride gas.

14.2 Nitric Acid

Nitric acid is prepared by the action of non-volatile sulphuric acid on any nitrate.

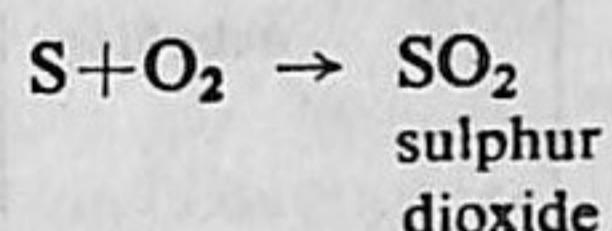
Experiment 2: Put a few grams of potassium nitrate in an all-glass retort and add concentrated sulphuric acid to it. Heat the mixture gently. The vapour of the acid is cooled and collected as an almost colourless liquid (Diagram 14.3). Potassium hydrogen sulphate is left in the flask.



Nitric acid thus obtained is slightly yellow. This is because of slight decomposition of the acid due to heating, when reddish brown fumes of nitrogen dioxide are given out which dissolve in the acid. Gentle heating is, therefore, essential in the preparation of the acid. Since nitric acid fumes attack rubber, an all-glass apparatus must be used.

14.3 Sulphuric Acid (The Contact Process)

First sulphur dioxide is prepared by burning sulphur or by heating strongly a sulphide mineral in excess of air.



Sulphur dioxide, on combining with oxygen forms sulphur trioxide, which on reacting with water gives sulphuric acid.

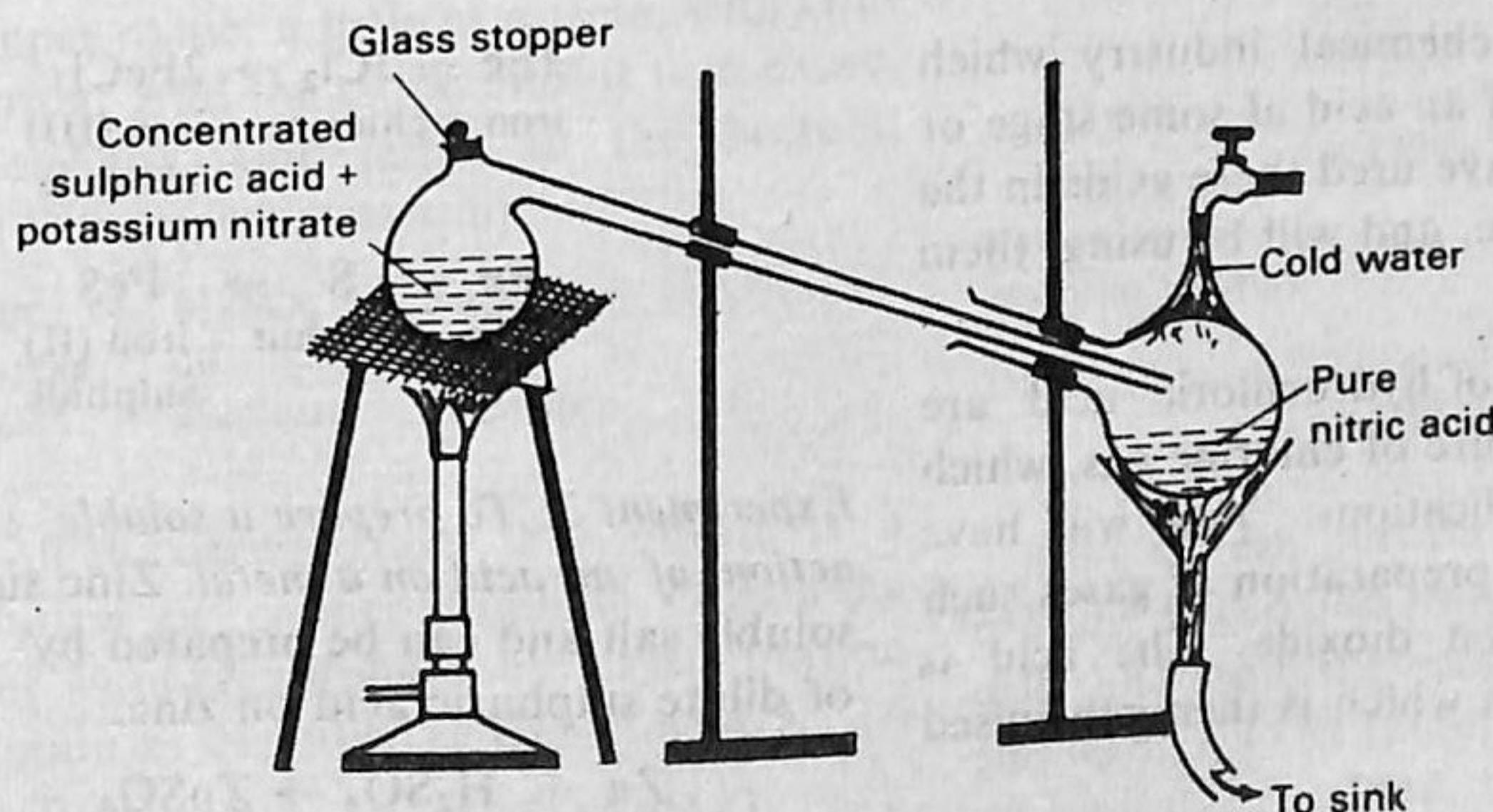
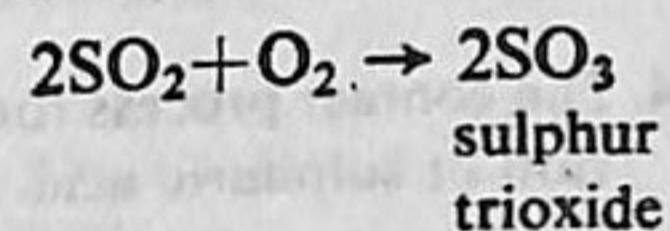
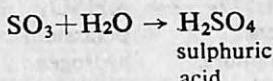


DIAGRAM 14.3 Preparation of nitric acid.



The combination between sulphur dioxide and oxygen takes place at a temperature of about 450°C in the presence of fine particles of platinum deposited on asbestos wool (Diagram 14.4). Platinum acts as a catalyst. Vanadium pentoxide is another catalyst now being used widely because it is cheaper.

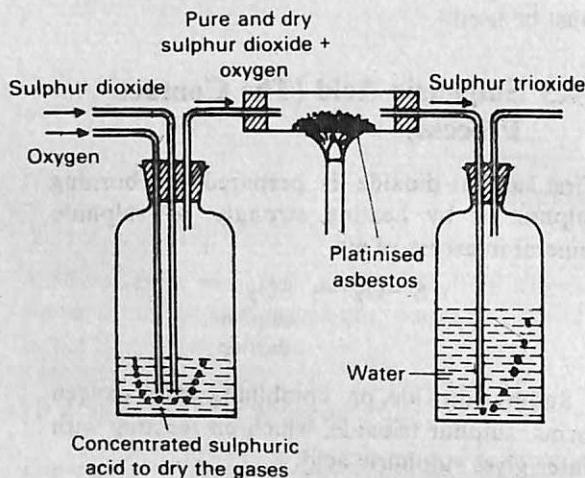


DIAGRAM 14.4 The contact process for the preparation of sulphuric acid.

14.4 Uses of Acids

There is hardly any chemical industry which does not make use of an acid at some stage or the other. You will have used these acids in the laboratory and outside, and will be using them a lot in future.

1. Large quantities of hydrochloric acid are used for the manufacture of chlorine gas, which has a number of applications. You will have used this acid in the preparation of gases such as hydrogen and carbon dioxide. The acid is used for cleaning iron which is then galvanised to check rusting.

2. Nitric acid is used in the manufacture of dyes, drugs, synthetic fibres and plastics. It

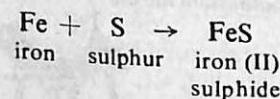
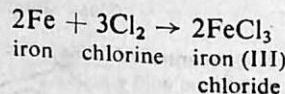
finds important use in the manufacture of fertilisers such as ammonium nitrate. Explosives such as nitrocellulose and T.N.T. are made from nitric acid.

3. Sulphuric acid is the backbone of chemical industries. So much so that the chemical wealth of a nation is said to be measured by how much sulphuric acid the nation uses. It is used in the manufacture of fertilisers such as ammonium sulphate and superphosphate of lime. It is used in preparing other acids, in refining of petroleum, in the preparation of detergents, synthetic fibres and pigments, etc.

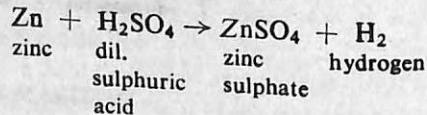
14.5 Preparation of Salts

There are several methods of preparing salts; some of the commonly used methods are given in the following.

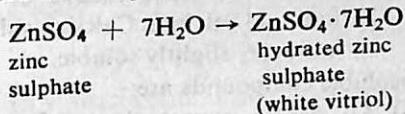
By synthesis: Some salts, such as iron (III) chloride and iron (II) sulphide, are prepared by combining elements present in the salts. When dry chlorine is passed over a heated iron wire, dark brown iron (III) chloride is formed. The reaction is highly exothermic and the glow of the iron continues even if the flame heating the iron wire is removed after some time. The formation of iron (II) sulphide is also an exothermic reaction.



Experiment 3: To prepare a soluble salt by the action of an acid on a metal. Zinc sulphate is a soluble salt and can be prepared by the action of dilute sulphuric acid on zinc.

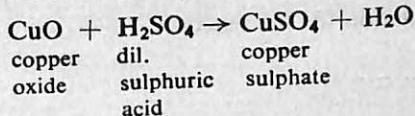


Take about 50 cm^3 of dilute sulphuric acid in a beaker and heat it on a wire gauze. Add small pieces of zinc, a little at a time, with constant stirring. You will observe effervescence because of the evolution of hydrogen gas. Add zinc till it is in excess and some of the metal settles at the base of the beaker. The acid has been used up completely. Filter off the metal and collect the filtrate in an evaporating dish. Evaporate the water by heating the filtrate till a saturated solution is formed and crystals start appearing. Cool when more crystals start appearing. The remaining liquid is drained off and the crystals are dried by filter paper. This process is called *crystallisation from solution*. The crystals are of hydrated zinc sulphate.

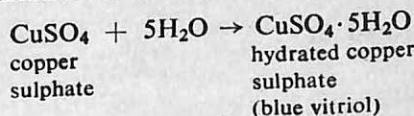


Experiment 4: To prepare a soluble salt by the action of acid on an insoluble oxide or hydroxide or carbonate. Copper sulphate is a soluble salt and can be prepared by the action of dilute sulphuric acid on copper oxide or copper carbonate, both of which are insoluble compounds.

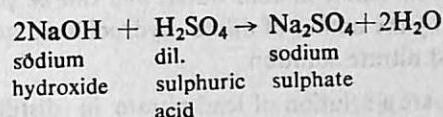
Take about 50 cm^3 of dilute sulphuric acid in a beaker and heat it on a wire gauze. Add the black copper oxide, a little at a time, with constant stirring. Add the oxide until it is in excess and some of the oxide settles at the bottom. The acid is completely used up.



Filter off the excess oxide. The filtrate is a blue solution of copper sulphate. Obtain the crystals of copper sulphate by crystallisation from solution as described in Experiment 3.



Experiment 5: To prepare a soluble salt by the action of an acid on a soluble oxide, hydroxide or carbonate. Sodium sulphate is a soluble salt and can be prepared by the action of dilute sulphuric acid on sodium hydroxide which is a soluble base.



Prepare solutions of sodium hydroxide and sulphuric acid in distilled water. Suck in 25 cm^3 of sodium hydroxide solution in a *pipette* and transfer this solution into a beaker. Add two drops of an *indicator*, say litmus solution. Fill the diluted sulphuric acid into a clean *burette* upto the zero mark. Run the acid into the alkali, very slowly, until one drop of the acid turns the litmus red. Read the volume of the acid needed for *neutralising* 25cm^3 of the alkali solution (Diagram 14.5). This process is known as *titration*.

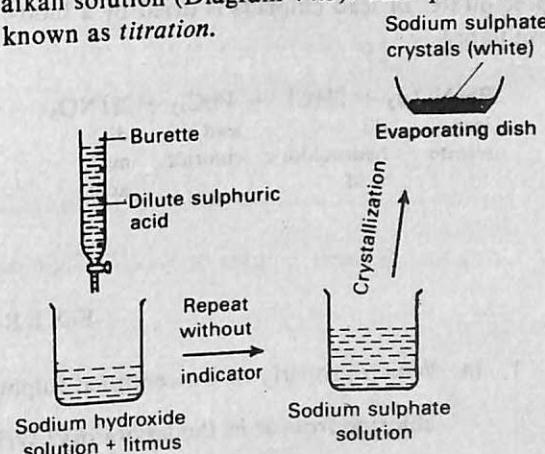
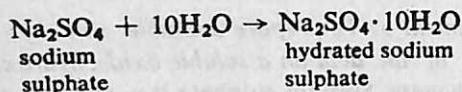


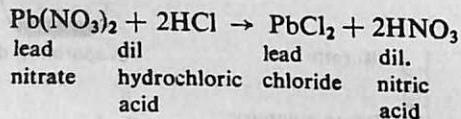
DIAGRAM 14.5 Titration of sodium hydroxide and sulphuric acid.

Take a fresh portion of 25 cm^3 of the same sodium hydroxide solution in an evaporating dish and add the same volume of the diluted sulphuric acid as found by titration; this time no indicator is added. Obtain crystals of sodium sulphate as described earlier by the method of crystallisation.



Experiment 6: To prepare an insoluble compound by a double decomposition reaction: Lead chloride is insoluble in cold water, and can be prepared by the action of dilute hydrochloric acid on lead nitrate solution.

Prepare a solution of lead nitrate in distilled water. Take about 50 cm^3 of this solution in a beaker and slowly add dilute hydrochloric acid, until further addition of the acid forms no precipitate. The heavy precipitate of lead chloride settles at the bottom of the beaker. Filter to get the precipitate on the filter paper. Wash the insoluble lead chloride on the filter paper several times with cold distilled water to remove any excess of acid sticking to the precipitate. Cold water and cold solutions are used because lead chloride is soluble in hot water. The white precipitate of lead chloride is dried by a blotting paper.



In case the insoluble substance does not dissolve on heating, warm solutions are used in the precipitation and hot water is used in washing the precipitate.

14.6 Solubility of Compounds

The following list gives the solubility of various compounds in water.

1. Soluble compounds are:
 - (a) All nitrates and hydrogen carbonates; all sodium, potassium and ammonium salts.
 - (b) All chlorides except those of silver and lead. Lead chloride is, however, soluble in hot water.
 - (c) All sulphates except those of lead, barium and calcium. Calcium sulphate is, however, slightly soluble.
2. Insoluble compounds are:
 - (a) All oxides except those of sodium and potassium.
 - (b) All hydroxides except those of sodium, potassium and ammonium. Calcium, magnesium, barium and lead hydroxides are slightly soluble.
 - (c) All carbonates except those of sodium, potassium and ammonium.

E X E R C I S E S

1. (a) Which property of concentrated sulphuric acid makes it possible to prepare hydrogen chloride from it in the laboratory? Write an equation for the reaction
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- (b) How is a solution of hydrogen chloride prepared in water?

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2. (a) How is nitric acid prepared in the laboratory? Draw the diagram of the apparatus used.....

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(b) Why is nitric acid thus prepared yellow?

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.....
.....

(c) How would you make the yellow acid colourless?.....

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.....
.....

(d) Mention two industrial uses of nitric acid

3. (a) State the conditions necessary for changing sulphur dioxide to sulphur trioxide.....

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.....
.....

(b) Write equations for preparing sulphuric acid from sulphur.....

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.....

(c) State three uses of sulphuric acid

4. (a) How would you prepare crystals of magnesium sulphate?

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(b) How would you prepare copper sulphate crystals from copper carbonate?

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(c) Write equations for the preparation of the two salts mentioned above

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5. (a) How is solid sodium chloride prepared from sodium carbonate? Write the equation for the reaction.....

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(b) How would you prepare calcium sulphate from calcium nitrate? Give the equation for the reaction.....

15

Metals and Alloys

Metals are of great importance to man. Metals such as iron, aluminium, copper and lead find various applications in our everyday life. There are many other metals we do not see around us, which are also important, e.g. uranium, a very rare metal, is the basic material for generation of atomic energy.

The metallic resources of a nation are of immense value. Silver and gold were the first metals discovered. The discovery of bronze, an alloy of copper and tin, resulted in the *Bronze Age* in history. This period was followed by the *Iron Age*, with the discovery of superior quality of iron. A number of uses of metals are based on their physical properties such as malleability, ductility, conduction of electricity and high tensile strength. These properties have been discussed in earlier chapters.

15.1 Occurrence of Metals

Metals which are not chemically reactive, e.g. gold and silver, are found in the uncombined state called the *native state*. This explains why gold and silver were discovered early.

Minerals of metals occur generally as oxides,

sulphides, carbonates and chlorides. These minerals are mixed with gravel, sand and other earthy impurities from which they must be separated by physical or chemical processes.

The minerals present in the earth's crust are decreasing as they are being used in extraction of metals. The National Institute of Oceanography has completed a survey discovering good polymetallic nodule fields in the Indian Ocean. It is expected that in years to come extraction of some useful metals from the nodules should be possible on a commercial scale.

15.2 Classification of Metals

There are many ways of classifying metals. Some of these are as follows.

1. *Classification based on reactivity:* You have learnt in an earlier chapter that metals differ in their reactivity. This can be shown by the reaction of metals with water and oxygen of the air. On this basis we can arrange the metals in a series called the *reactivity series*. This series has been given earlier.

2. *Classification based on other properties:* Metals can be divided into two groups.

(a) Iron and steel which form the backbone of industries.

(b) Non-ferrous metals which can be further subdivided into the following groups.

(i) Light metals, e.g. potassium, sodium, magnesium and aluminium.

(ii) Heavy metals, e.g. lead, mercury, copper, silver, gold and platinum.

(iii) Protective metals, e.g. chromium, nickel, zinc and tin.

15.3 Extraction of Metals

Minerals are compounds of metals found in the earth's crust. The mixture of minerals and the earthy impurities such as sand and gravel form the ores of the metals. Some of the principles for the extraction of metals from their ores are as follows.

1. Reactive metals such as potassium, sodium, calcium and magnesium are extracted by the *electrolysis of their molten chlorides*.

For example, sodium chloride whose melting point is 800°C is mixed with calcium chloride which brings the melting point of the former to about 600°C . When electricity is passed through the molten mixture of the two salts, sodium, being a positively charged ion, is collected around a steel cathode. Chloride ions are discharged on a graphite anode and form molecules of chlorine which are collected as greenish-yellow chlorine gas. Thus:

At the cathode: Sodium ions change to sodium atoms.

At the anode: Chloride ions change to chlorine gas.

2. Less reactive metals are extracted by the method which generally involves the following steps.

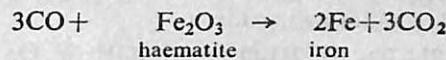
(a) *Concentration of the ore:* This step leads to the removal of earthy impurities.

(b) *Roasting of the ore:* The purpose of roasting, or heating in air, is to change carbonate and sulphide ores of metals to oxides.

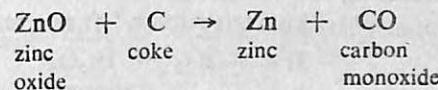
(c) *Reduction:* The oxide is reduced to the metal by removal of oxygen. The common reducing agent used is coke, a form of carbon. The reduction is carried out in a furnace whose temperature is about $1000\text{-}1500^{\circ}\text{C}$. For example, the ore of iron called haematite is mixed with coke and heated in a blast furnace. The coke first burns in the furnace to produce carbon dioxide which changes to carbon monoxide.



The carbon monoxide reduces the haematite to form iron.



The ore of zinc when roasted in air changes to zinc oxide, which is mixed with coke and heated to give the metal.



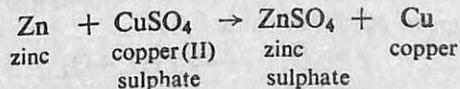
(d) *Refining of metal:* This is often done by electrolysis but other processes are also in use.

15.4 Chemical Properties of Metals

Some of the chemical properties of metals have already been discussed earlier. The properties are related to reactivity series.

1. As we go down in the reactivity series the metals react with water at higher temperatures. Copper and other metals below hydrogen do not react with water. Aluminium, even though a reactive metal, does not react with water because of its compact and rigid oxide layer.

2. A more reactive metal displaces a less reactive metal from its compound.



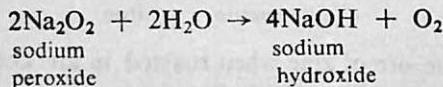
3. Dilute hydrochloric and sulphuric acids react with metals above hydrogen in the reactivity series. Hence these acids have no reaction with copper, silver and mercury.

4. Metals react with oxygen to form basic oxides. Reactive metals combine at a lower temperature.

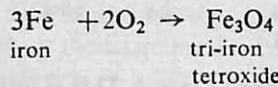
(a) Sodium when warmed burns with a yellow flame in oxygen to form a white solid. This is a higher oxide and contains more oxygen than present in the normal oxide of the metal. It is called sodium peroxide.



Sodium peroxide reacts with water to give the alkali, sodium hydroxide.

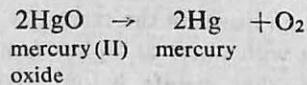


(b) Red hot steel wool also burns in oxygen producing showers of sparks. A black oxide of iron, call tri-iron tetroxide is left as a residue.

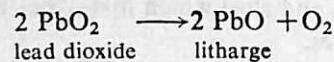
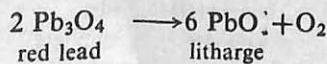


Other metals combine with oxygen to form their normal oxides. Lead combines at 850°C to form its normal oxide, called lead monoxide, on its surface.

5. (a) Oxides of metals do not decompose on heating except oxides of silver and mercury which produce the metal, liberating oxygen.



(b) Lead forms two other oxides, besides the normal lead (II) oxide. Red lead, Pb_3O_4 , is a red powder and lead dioxide, PbO_2 , is a chocolate brown powder. When these are heated they change into yellow lead monoxide, also called litharge, which fuses with glass. Oxygen is given off at the same time.



15.5 Uses of some Metals

1. The impure iron, called cast iron, is used for castings, water pipes and railings. Steel is used in making all sorts of things from small tools to warships and military tanks. It is the backbone of industries.

2. Copper and aluminium are good conductors of electricity. Hence they are used in electrical transmission. Copper, being non-reactive, is used in making boiler pipes. Aluminium, mixed with some other metals, is used in making aircrafts.

3. Lead is used in making water pipes. Car batteries also contain lead. Lead is used as a protective against radioactive radiations.

4. Zinc is used in galvanising iron and in dry cells.

15.6 Alloys

Alloys are mixtures of metals, prepared by melting and mixing the component metals in definite proportions by weight.

Alloys are prepared to obtain properties which are not possessed by component metals. The properties which change in alloying metals are hardness, malleability, resistance to *corrosion* (which means chemical attack by atmosphere, water and soil) and melting point. *Amalgam* is an alloy of a metal with mercury.

The composition, properties and uses of some of the common alloys are given in Table 15.1.

PRACTICAL WORK

Identification of Ions

(a) Non-metallic ions

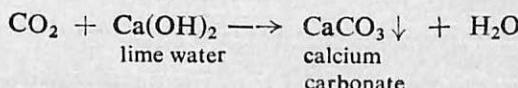
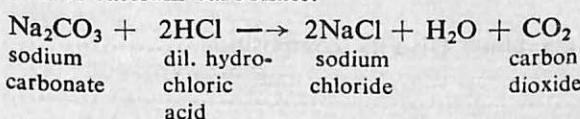
The non-metallic ions generally present in

TABLE 15.1 Composition, properties and uses of common alloys

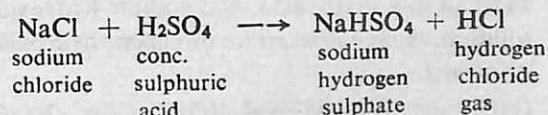
Name of the alloy	Composition	Use	Properties
1. Brass	Copper, zinc.	Cartridge cases, musical instruments, machine parts.	Easily moulded, resists corrosion.
2. Bronze	Copper, tin.	Statues, coins, medals.	Hard, can be easily cast.
3. Duralumin	Aluminium, copper, magnesium and manganese.	Aircraft.	Light and tough.
4. Magnalium	Aluminium, magnesium.	Scientific instruments, aircraft.	Light, tough and resists corrosion.
5. German silver	Copper, zinc, nickel.	Utensils and decorative pieces.	Silvery appearance and resists corrosion.
6. Solder	Lead, tin.	Soldering joints, fuse.	Low melting point, alloys with soldered surfaces.
7. Type metal	Lead, tin, antimony.	Type used in printing.	Expands on solidification.
8. Stainless steel	Iron, chromium, nickel.	Cutlery and machine parts.	Resists corrosion.
9. Pewter	Tin, copper, bismuth and antimony.	Carburettor, table ware.	Cast easily and hard.

salts are carbonates, chlorides, sulphates and nitrates.

(i) **Carbonate:** When dilute hydrochloric acid is added to a metallic carbonate, there is effervescence due to production of carbon dioxide gas. The gas is colourless and odourless and turns lime water milky due to white insoluble particles of calcium carbonate.

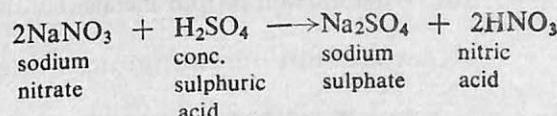


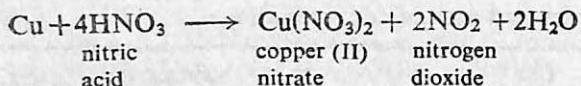
(ii) **Chloride:** The salt containing a chloride ion when treated with concentrated sulphuric acid produces hydrogen chloride gas.



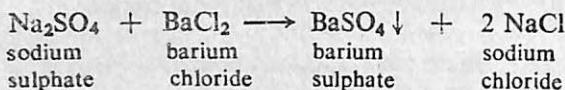
The hydrogen chloride gas fumes in air. It produces dense white fumes with a rod dipped in ammonia solution. The gas gives a white precipitate with silver nitrate solution in dilute nitric acid.

(iii) **Nitrate:** The salt when heated with concentrated sulphuric acid and two or three pieces of copper turnings, produces dense reddish brown fumes of nitrogen dioxide.





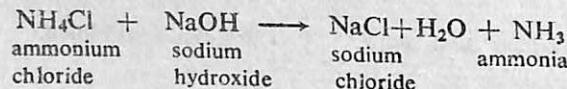
(iv) *Sulphate*: Prepare the salt solution in dilute hydrochloric acid and add barium chloride solution. A white precipitate of barium sulphate is produced which is insoluble in concentrated hydrochloric or nitric acid.



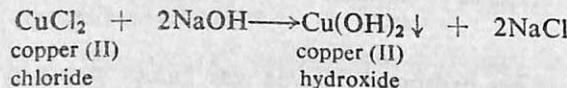
(b) Metallic ions

Sodium hydroxide solution is used to identify many metallic ions.

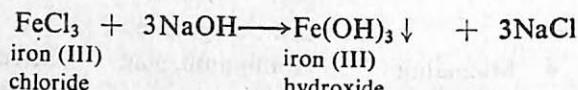
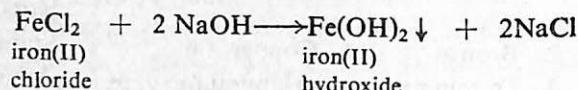
(i) *Ammonium*: To the salt add sodium hydroxide solution and heat. A pungent gas smelling like ammonium hydroxide is given off. The gas is ammonia and turns red litmus blue. With a rod dipped in concentrated hydrochloric acid, dense white fumes are produced.



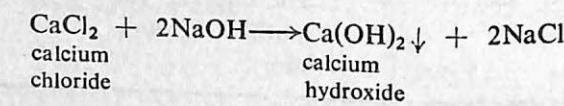
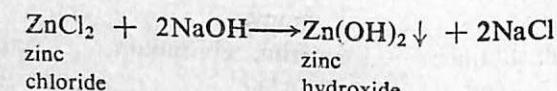
(ii) *Copper*: To the solution of the salt in water or in a dilute acid, add sodium hydroxide solution. A *blue precipitate* of copper hydroxide is formed.



(iii) *Iron(II) and iron(III)*: To the salt solution in water or in a dilute acid add sodium hydroxide solution. Iron(II) salts give *dirty green precipitate* of iron(II) hydroxide and iron(III) salts give a *reddish brown precipitate* of iron(III) hydroxide.



(iv) *Zinc and calcium*: To the salt solution in water or in a dilute acid, when sodium hydroxide solution is added, a *white precipitate* indicates the presence of either zinc or calcium salts.



The white precipitate of zinc hydroxide dissolves in excess of sodium hydroxide solution, whereas the white precipitate of calcium hydroxide solution does not dissolve in excess of sodium hydroxide solution.

EXERCISES

1. (a) Which was the earliest alloy known to mankind? Give its composition.....
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- (b) What are non ferrous metals?.....
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(c) Describe the method involved in extraction of sodium metal.....

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2. (a) The metal zinc is extracted by roasting its sulphide ore followed by reduction. Write equations for the reactions taking place.....

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(b) Iron is extracted by putting coke and its oxide ore in a blast furnace. Write equations for the reactions taking place in the furnace.....

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.....
.....

3. (a) How can the metal calcium be extracted?.....

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.....

(b) Why is aluminium much cheaper today than it was hundred years ago?

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.....

4. What would you observe when the following are heated in air? Write an equation for each reaction.

(a) Mercury oxide

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.....

(b) Red lead.....

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.....

(c) Sodium metal.....
.....
.....

(d) Iron filings

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.....

5. Give two uses of the metals mentioned below. Give reasons for each use.

(a) Copper.....
.....

(b) Aluminium.....
.....

(c) Lead.....
.....

(d) Zinc.....
.....

6. (a) Define an alloy.....
.....

(b) Give the composition and uses of the following alloys, mentioning the reason for each use.

(i) Brass

.....
.....

(ii) Duralumin.....
.....

(iii) Magnalium.....
.....

(iv) Solder

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7. How would you distinguish between the following pairs of compounds?

(a) Sodium chloride and sodium nitrate.....

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(b) Copper sulphate and ammonium sulphate.....

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(c) Zinc nitrate and calcium nitrate

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8. A blue crystalline compound A when heated gives water leaving a white solid residue B. The solution of the compound A when treated with sodium hydroxide solution gives a blue precipitate of the substance C, which when heated gives a black compound D.

(a) Name the substances A, B, C and D.

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(b) Write equations for the reactions taking place in the experiment mentioned above.....

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9. A pale green compound A is dissolved in water. When the solution is treated with sodium hydroxide solution, a dirty green precipitate of compound B is formed. When the aqueous solution of compound A is treated with concentrated nitric acid, a yellow solution is

obtained. On reacting with sodium hydroxide this solution gives a reddish brown precipitate of substance C.

(a) Name the compounds A, B and C

.....

(b) What change takes place when the compound A is treated with concentrated nitric acid?

.....

.....

(c) Mention the type of change taking place when nitric acid reacts with compound A.....

.....

10. A white crystalline solid A sublimes on heating. When the compound A is heated with sodium hydroxide solution, a pungent gas B is given out which turns red litmus blue. The gas B when reacted with hydrogen chloride gas produces the white crystalline solid A.

(a) Name the substances A and B.....

(b) Write equations for the following reactions.

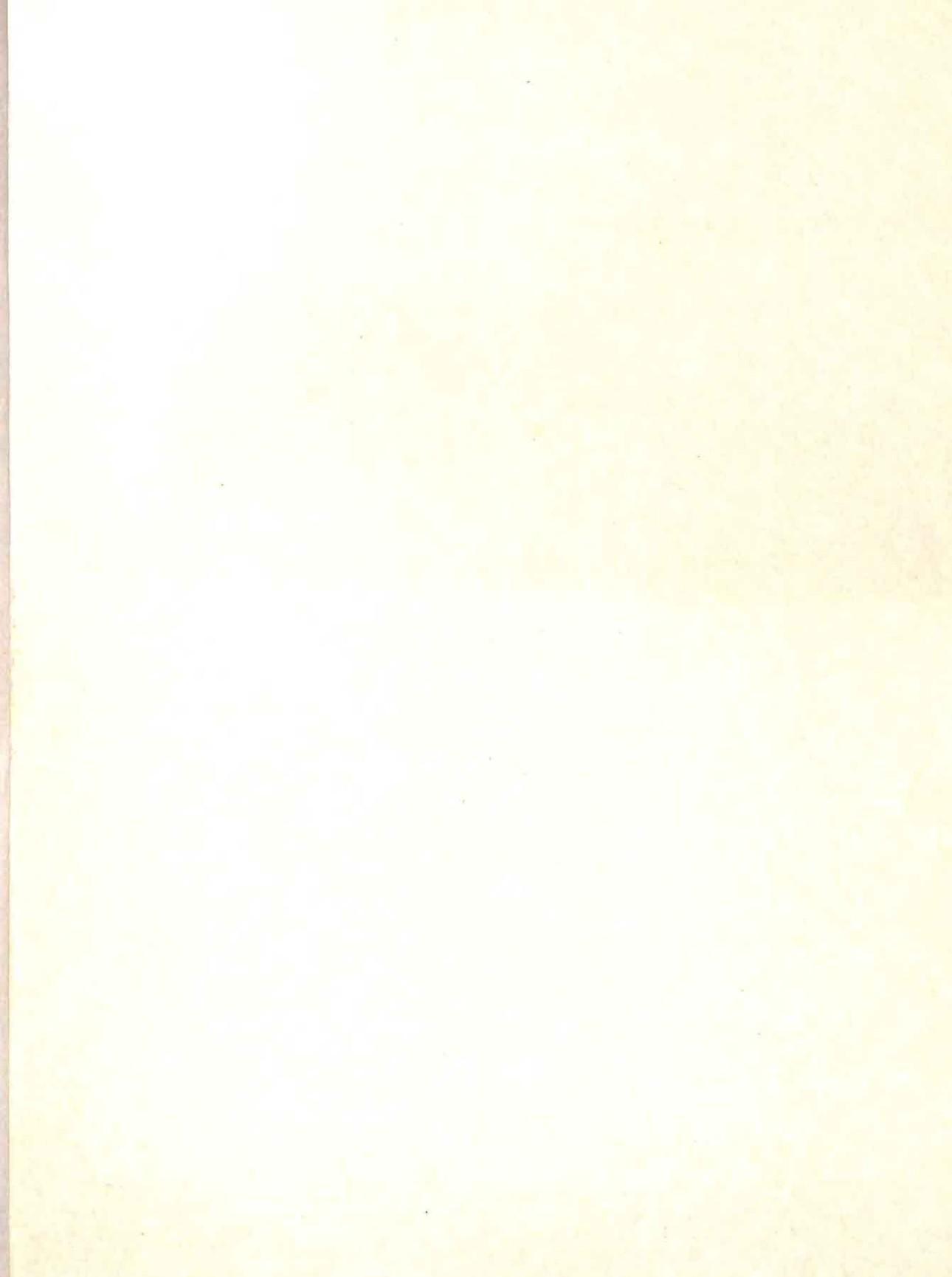
(i) When solid A is heated.....

(ii) When solid A is heated with sodium hydroxide solution

.....

(iii) When gas B reacts with hydrogen chloride.....

.....



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